Los Alamos National Laboratory's Chemistry Division Presents

Periodic Table of the Elements

A Resource for Elementary, Middle School, and High School Students

Click an element for more information:

Group**

Period	1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8	9 ------- VIII ------- ------- 8 -------	10	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 vIIIA 8A
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.47	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 190.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.5	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (210)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac~ (227)	104 Rf (257)	105 Db (260)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 --- ()	111 --- ()	112 --- ()		114 --- ()		116 --- ()		118 --- ()

Lanthanide Series*	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (147)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
Actinide Series~	90 Th 232.0	91 Pa (231)	92 U (238)	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (249)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (254)	103 Lr (257)

** Groups are noted by 3 notation conventions.

For a list of a the element names and symbols in alphabetical order, click here

---

Have a question - comment - suggestion,
send us feedback or email to cstis@lanl.gov

What is the Periodic Table?

How to use the Periodic Table

Click here to see Mendeleev's original Periodic Table

Chemistry in a Nutshell

Naming New Elements

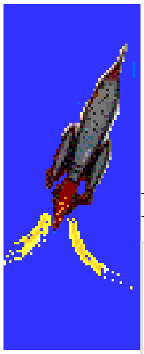
LANL DOE Disclaimer

Number of visitors since May 2000: 621646

Last Updated: 08/29/00

about this resource

Periodic Table



Hydrogen

For rocket fuel

Atomic Number:	1
Atomic Symbol:	H
Atomic Weight:	1.0079
Electron Configuration:	$1s^1$

History

(Gr. hydro, water, and genes, forming) Hydrogen was prepared many years before it was recognized as a distinct substance by Cavendish in 1776.

Named by Lavoisier, hydrogen is the most abundant of all elements in the universe. The heavier elements were originally made from Hydrogen or from other elements that were originally made from Hydrogen.

Sources

Hydrogen is estimated to make up more than 90% of all the atoms or three quarters of the mass of the universe. This element is found in the stars, and plays an important part in powering the universe through both the proton-proton reaction and carbon-nitrogen cycle -- stellar hydrogen fusion processes that release massive amounts of energy by combining Hydrogen to form Helium.

Production of hydrogen in the U.S. alone now amounts to about 3 billion cubic feet per year. Hydrogen is prepared by

- steam on heated carbon,
- decomposition of certain hydrocarbons with heat,
- action of sodium or potassium hydroxide on aluminum
- electrolysis of water, or
- displacement from acids by certain metals.

Liquid hydrogen is important in cryogenics and in the study of superconductivity, as its melting point is only 20 degrees above absolute zero.

Tritium is readily produced in nuclear reactors and is used in the production of the hydrogen bomb.

Hydrogen is the primary component of Jupiter and the other gas giant planets. At some depth in the

planet's interior the pressure is so great that solid molecular hydrogen is converted to solid metallic hydrogen.

In 1973, a group of Russian experimenters may have produced metallic hydrogen at a pressure of 2.8 Mbar. At the transition the density changed from 1.08 to 1.3 g/cm3. Earlier, in 1972, at Livermore, California, a group also reported on a similar experiment in which they observed a pressure-volume point centered at 2 Mbar. Predictions say that metallic hydrogen may be metastable; others have predicted it would be a superconductor at room temperature.

Compounds

Although pure Hydrogen is a gas we find very little of it in our atmosphere. Hydrogen gas is so light that uncombined Hydrogen will gain enough velocity from collisions with other gases that they will quickly be ejected from the atmosphere. On earth, hydrogen occurs chiefly in combination with oxygen in water, but it is also present in organic matter such as living plants, petroleum, coal, etc. It is present as the free element in the atmosphere, but only to the extent of less than 1 ppm by volume. The lightest of all gases, hydrogen combines with other elements -- sometimes explosively -- to form compounds.

Uses

Great quantities are required commercially for the fixation of nitrogen from the air in the Haber ammonia process and for the hydrogenation of fats and oils. It is also used in large quantities in methanol production, in hydrodealkylation, hydrocracking, and hydrodesulfurization. Other uses include rocket fuel, welding, producing hydrochloric acid, reducing metallic ores, and filling balloons.

The lifting power of 1 cubic foot of hydrogen gas is about 0.07 lb at 0C, 760 mm pressure.

The Hydrogen Fuel cell is a developing technology that will allow great amounts of electrical power to be obtained using a source of hyrogen gas.

Consideration is being given to an entire economy based on solar- and nuclear-generated hydrogen. Public acceptance, high capital investment, and the high cost of hydrogen with respect to today's fuels are but a few of the problems facing such an economy. Located in remote regions, power plants would electrolyze seawater; the hydrogen produced would travel to distant cities by pipelines. Pollution-free hydrogen could replace natural gas, gasoline, etc., and could serve as a reducing agent in metallurgy, chemical processing, refining, etc. It could also be used to convert trash into methane and ethylene.

Forms

Quite apart from isotopes, it has been shown that under ordinary conditions hydrogen gas is a mixture of two kinds of molecules, known as ortho- and para-hydrogen, which differ from one another by the spins of their electrons and nuclei.

Normal hydrogen at room temperature contains 25% of the para form and 75% of the ortho form. The ortho form cannot be prepared in the pure state. Since the two forms differ in energy, the physical properties also differ. The melting and boiling points of parahydrogen are about 0.1$_o$C lower than those of normal hydrogen.

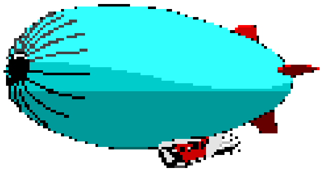
Isotopes

The ordinary isotope of hydrogen, H, is known as Protium, the other two isotopes are Deuterium (a proton and a neutron) and Tritium (a protron and two neutrons). Hydrogen is the only element whose isotopes have been given different names. Deuterium and Tritium are both used as fuel in nuclear fusion reactors. One atom of Deuterium is found in about 6000 ordinary hydrogen atoms.

Deuterium is used as a moderator to slow down neutrons. Tritium atoms are also present but in much smaller proportions. Tritium is readily produced in nuclear reactors and is used in the production of the hydrogen (fusion) bomb. It is also used as a radioactive agent in making luminous paints, and as a tracer.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team



Helium

For blimps

Atomic Number:	2
Atomic Symbol:	He
Atomic Weight:	4.00260
Electron Configuration:	$1s^2$

History

(Gr. helios, the sun). Janssen obtained the first evidence of helium during the solar eclipse of 1868 when he detected a new line in the solar spectrum. Lockyer and Frankland suggested the name helium for the new element. In 1895 Ramsay discovered helium in the uranium mineral clevite while it was independently discovered in cleveite by the Swedish chemists Cleve and Langlet at about the same time. Rutherford and Royds in 1907 demonstrated that alpha particles are helium nuclei.

Sources

Except for hydrogen, helium is the most abundant element found through out the universe. Helium is extracted from natural gas. In fact, all natural gas contains at least trace quantities of helium.

It has been detected spectroscopically in great abundance, especially in the hotter stars, and it is an important component in both the proton-proton reaction and the carbon cycle, which account for the energy of the sun and stars.

The fusion of hydrogen into helium provides the energy of the hydrogen bomb. The helium content of the atmosphere is about 1 part in 200,000. While it is present in various radioactive minerals as a decay product, the bulk of the Free World's supply is obtained from wells in Texas, Oklahoma, and Kansas. The only known helium extraction plants, outside the United States, in 1984 were in Eastern Europe (Poland), the USSR, and a few in India.

Cost

The cost of helium fell from $2500 / ft^{3} in 1915 to 1.5 cents / ft^{3} in 1940. The U.S. Bureau of Mines has set the price of Grade A helium at $37.50 / 1000 ft^{3} in 1986.

Properties

Helium has the lowest melting point of any element and is widely used in cryogenic research because its boiling point is close to absolute zero. Also, the element is vital in the study of super conductivity.

Using liquid helium, Kurti and co-workers and others, have succeeded in obtaining temperatures of a few microkelvins by the adiabatic demagnetization of copper nuclei.

It has other peculiar properties. Helium is the only liquid that cannot be solidified by lowering the temperature. It remains liquid down to absolute zero at ordinary pressures, but it can readily be solidified by increasing the pressure. Solid 3He and 4He are unusual in that both can be changed in volume by more than 30% by applying pressure.

The specific heat of helium gas is unusually high. The density of helium vapor at the normal boiling point is also very high, with the vapor expanding greatly when heated to room temperature. Containers filled with helium gas at 5 to 10 K should be treated as though they contained liquid helium due to the large increase in pressure resulting from warming the gas to room temperature.

While helium normally has a 0 valence, it seems to have a weak tendency to combine with certain other elements. Means of preparing helium difluoride have been studied, and species such as HeNe and the molecular ions He+ and He++ have been investigated.

Isotopes

Seven isotopes of helium are known: Liquid helium (He4) exists in two forms: He4I and He4II, with a sharp transition point at 2.174K. He4I (above this temperature) is a normal liquid, but He4II (below it) is unlike any other known substance. It expands on cooling; its conductivity for heat is enormous; and neither its heat conduction nor viscosity obeys normal rules.

Uses

- as an inert gas shield for arc welding;
- a protective gas in growing silicon and germanium crystals and producing titanium and zirconium;
- as a cooling medium for nuclear reactors, and
- as a gas for supersonic wind tunnels.

A mixture of helium and oxygen is used as an artificial atmosphere for divers and others working under pressure. Different ratios of He/O2 are used for different depths at which the diver is operating.

Helium is extensively used for filling balloons as it is a much safer gas than hydrogen. One of the recent largest uses for helium has been for pressuring liquid fuel rockets. A Saturn booster, like the type used on the Apollo lunar missions, required about 13 million ft^3 of helium for a firing, plus more for checkouts.

Liquid helium's use in magnetic resonance imaging (MRI) continues to increase as the medical profession accepts and develops new uses for the equipment. This equipment has eliminated some need for exploratory surgery by accurately diagnosing patients. Another medical application uses MRE to determine (by blood analysis) whether a patient has any form of cancer.

Helium is also being used to advertise on blimps for various companies, including Goodyear. Other lifting gas applications are being developed by the Navy and Air Force to detect low-flying cruise missiles. Additionally, the Drug Enforcement Agency is using radar-equipped blimps to detect drug smugglers along the United States boarders. In addition, NASA is currently using helium-filled balloons to sample the atmosphere in Antarctica to determine what is depleting the ozone layer.

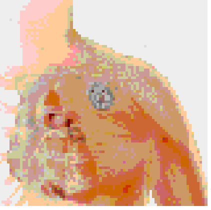
Costs

Materials which become super conductive at higher temperatures than the boiling point of helium could have a major impact on the demand for helium. These less costly refrigerant materials could replace the present need to cool superconductive materials to the boiling point of helium.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Lithium

For pacemaker batteries

Atomic Number:	3
Atomic Symbol:	Li
Atomic Weight:	6.941
Electron Configuration:	$[He]2s^1$

History

(Gr. lithos, stone) Discovered by Arfvedson in 1817. Lithium is the lightest of all metals, with a density only about half that of water.

Sources

It does not occur free in nature; combined it is found in small units in nearly all igneous rocks and in the waters of many mineral springs. Lepidolite, spodumeme, petalite, and amblygonite are the more important minerals containing it.

Lithium is presently being recovered from brines of Searles Lake, in California, and from those in Nevada. Large deposits of quadramene are found in North Carolina. The metal is produced electrolytically from the fused chloride. Lithium is silvery in appearance, much like Na and K, other members of the alkali metal series. It reacts with water, but not as vigorously as sodium. Lithium imparts a beautiful crimson color to a flame, but when the metal burns strongly, the flame is a dazzling white.

Uses

Since World War II, the production of lithium metal and its compounds has increased greatly. Because the metal has the highest specific heat of any solid element, it has found use in heat transfer applications; however, it is corrosive and requires special handling. The metal has been used as an alloying agent, is of interest in synthesis of organic compounds, and has nuclear applications. It ranks as a leading contender as a battery anode material as it has a high electrochemical potential. Lithium is used in special glasses and ceramics. The glass for the 200-inch telescope at Mt. Palomar contains lithium as a minor ingredient. Lithium chloride is one of the most lyproscopic materials known, and it, as well as lithium bromide, is used in air conditioning and industrial drying systems. Lithium stearate is used as an all-purpose and high-temperature lubricant. Other lithium compounds are used in dry cells and storage

batteries.

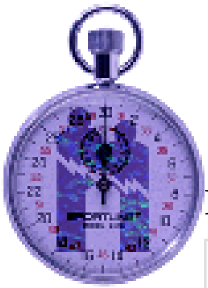
Cost

The metal is priced at about $300 / lb.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Beryllium

For watch springs

Atomic Number:	4
Atomic Symbol:	Be
Atomic Weight:	9.01218
Electron Configuration:	$[He]2s^2$

History

(Gr. beryllos, beryl; also called Glucinium or Glucinum, Gr. glykys, sweet) Discovered as the oxide by Vauquelin in beryl and in emeralds in 1798. The metal was isolated in 1828 by Wohler and by Bussy independently by the action of potassium on beryllium chloride.

Sources

Beryllium is found in some 30 mineral species, the most important of which are bertrandite, beryl, chrysoberyl, and phenacite. Aquamarine and emerald are precious forms of beryl. Beryl and bertrandite are the most important commercial sources of the element and its compounds. Most of the metal is now prepared by reducing beryllium fluoride with magnesium metal. Beryllium metal did not become readily available to industry until 1957.

Properties

The metal, steel gray in color, has many desirable properties. As one of the lightest of all metals, it has one of the highest melting points of the light metals. Its modulus of elasticity is about one third greater than that of steel. It resists attack by concentrated nitric acid, has excellent thermal conductivity, and is nonmagnetic. It has a high permeability to X-rays and when bombarded by alpha particles, as from radium or polonium, neutrons are produced in the amount of about 30 neutrons / million alpha particles.

At ordinary temperatures, beryllium resists oxidation in air, although its ability to scratch glass is probably due to the formation of a thin layer of the oxide.

Uses

Beryllium is used as an alloying agent in producing beryllium copper, which is extensively used for springs, electrical contacts, spot-welding electrodes, and non-sparking tools. It is applied as a structural material for high-speed aircraft, missiles, spacecraft, and communication satellites. Other uses include windshield frame, brake discs, support beams, and other structural components of the space shuttle.

Because beryllium is relatively transparent to X-rays, ultra-thin Be-foil is finding use in X-ray lithography for reproduction of microminiature integrated circuits.

Beryllium is used in nuclear reactors as a reflector or moderator for it has a low thermal neutron absorption cross section.

It is used in gyroscopes, computer parts, and instruments where lightness, stiffness, and dimensional stability are required. The oxide has a very high melting point and is also used in nuclear work and ceramic applications.

Handling

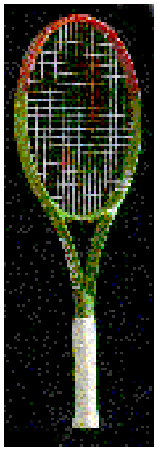
Beryllium and its salts are toxic and should be handled with the greatest of care. Beryllium and its compounds should not be tasted to verify the sweetish nature of beryllium (as did early experimenters). The metal, its alloys, and its salts can be handled if certain work codes are observed, but no attempt should be made to work with beryllium before becoming familiar with proper safeguards.

Isotopes available at Los Alamos National Laboratory

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Boron

For tennis rackets

Atomic Number:	5
Atomic Symbol:	B
Atomic Weight:	10.81
Electron Configuration:	$[He]2s^2 2p^1$

History

(Ar. Buraq, Pers. Burah) Boron compounds have been known for thousands of years, but the element was not discovered until 1808 by Sir Humphry Davy and by Gay-Lussac and Thenard.

Sources

The element is not found free in nature, but occurs as orthoboric acid usually found in certain volcanic spring waters and as borates in boron and colemantie. Ulexite, another boron mineral, is interesting as it is nature's own version of "fiber optics."

Important sources of boron are ore rasorite (kernite) and tincal (borax ore). Both of these ores are found in the Mojave Desert. Tincal is the most important source of boron from the Mojave. Extensive borax deposits are also found in Turkey.

Boron exists naturally as 19.78% 10B isotope and 80.22% 11B isotope. High-purity crystalline boron may be prepared by the vapor phase reduction of boron trichloride or tribromide with hydrogen on electrically heated filaments. The impure or amorphous, boron, a brownish-black powder, can be obtained by heating the trioxide with magnesium powder.

Boron of 99.9999% purity has been produced and is available commercially. Elemental boron has an energy band gap of 1.50 to 1.56 eV, which is higher than that of either silicon or germanium.

Properties

Optical characteristics include transmitting portions of the infrared. Boron is a poor conductor of electricity at room temperature but a good conductor at high temperature.

Uses

Amorphous boron is used in pyrotechnic flares to provide a distinctive green color, and in rockets as an igniter.

By far the most commercially important boron compound in terms of dollar sales is $Na_2B_4O_7.5H_2O$. This pentahydrate is used in very large quantities in the manufacture of insulation fiberglass and sodium perborate bleach.

Boric acid is also an important boron compound with major markets in textile products. Use of borax as a mild antiseptic is minor in terms of dollars and tons. Boron compounds are also extensively used in the manufacture of borosilicate glasses. Other boron compounds show promise in treating arthritis.

The isotope boron-10 is used as a control for nuclear reactors, as a shield for nuclear radiation, and in instruments used for detecting neutrons. Boron nitride has remarkable properties and can be used to make a material as hard as diamond. The nitride also behaves like an electrical insulator but conducts heat like a metal.

It also has lubricating properties similar to graphite. The hydrides are easily oxidized with considerable energy liberation, and have been studied for use as rocket fuels. Demand is increasing for boron filaments, a high-strength, lightweight material chiefly employed for advanced aerospace structures.

Boron is similar to carbon in that it has a capacity to form stable covalently bonded molecular networks. Carbonates, metalloboranes, phosphacarboranes, and other families comprise thousands of compounds.

Costs

Crystalline boron (99%) costs about $5/g. Amorphous boron costs about $2/g.

Handling

Elemental boron and the borates are not considered to be toxic, and they do not require special care in handling. However, some of the more exotic boron hydrogen compounds are definitely toxic and do require care.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Carbon

For pencils

Atomic Number:	6
Atomic Symbol:	C
Atomic Weight:	12.011
Electron Configuration:	[He]$2s^2 2p^2$

History

(Latin: carbo, charcoal) Carbon, an element of prehistoric discovery, is very widely distributed in nature. It is found in abundance in the sun, stars, comets, and atmospheres of most planets. Carbon in the form of microscopic diamonds is found in some meteorites.

Natural diamonds are found in kimberlite of ancient volcanic "pipes," found in South Africa, Arkansas, and elsewhere. Diamonds are now also being recovered from the ocean floor off the Cape of Good Hope. About 30% of all industrial diamonds used in the U.S. are now made synthetically.

The energy of the sun and stars can be attributed at least in part to the well-known carbon-nitrogen cycle.

Forms

Carbon is found free in nature in three allotropic forms: amorphous, graphite, and diamond. A fourth form, known as "white" carbon, is now thought to exist. Ceraphite is one of the softest known materials while diamond is one of the hardest.

Graphite exists in two forms: alpha and beta. These have identical physical properties, except for their crystal structure. Naturally occurring graphites are reported to contain as much as 30% of the rhombohedral (beta) form, whereas synthetic materials contain only the alpha form. The hexagonal alpha type can be converted to the beta by mechanical treatment, and the beta form reverts to the alpha on heating it above 1000$_{o}$C.

In 1969 a new allotropic form of carbon was produced during the sublimation of pyrolytic graphite at low pressures. Under free-vaporization conditions above ~2550$_{o}$K, "white" carbon forms as small transparent crystals on the edges of the planes of graphite. The interplanar spacings of "white" carbon are identical to those of carbon form noted in the graphite gneiss from the Ries (meteroritic) Crater of Germany. "White" carbon is a transparent birefringent material. Little information is presently available about this allotrope.

Compounds

In combination, carbon is found as carbon dioxide in the atmosphere of the earth and dissolved in all natural waters. It is a component of great rock masses in the form of carbonates of calcium (limestone), magnesium, and iron. Coal, petroleum, and natural gas are chiefly hydrocarbons.

Carbon is unique among the elements in the vast number and variety of compounds it can form. With hydrogen, oxygen, nitrogen, and other elements, it forms a very large number of compounds, carbon atom often being linked to carbon atom. There are close to ten million known carbon compounds, many thousands of which are vital to organic and life processes.

Without carbon, the basis for life would be impossible. While it has been thought that silicon might take the place of carbon in forming a host of similar compounds, it is now not possible to form stable compounds with very long chains of silicon atoms. The atmosphere of Mars contains 96.2% CO_2. Some of the most important compounds of carbon are carbon dioxide (CO_2), carbon monoxide (CO), carbon disulfide (CS_2), chloroform ($CHCl_3$), carbon tetrachloride (CCl_4), methane (CH_4), ethylene (C_2H_4), acetylene (C_2H_2), benzene (C_6H_6), acetic acid (CH_3COOH), and their derivatives.

Isotopes

Carbon has seven isotopes. In 1961 the International Union of Pure and Applied Chemistry adopted the isotope carbon-12 as the basis for atomic weights. Carbon-14, an isotope with a half-life of 5715 years, has been widely used to date such materials as wood, archaeological specimens, etc.

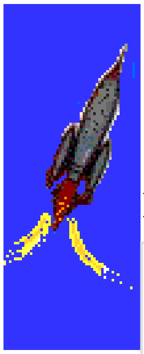
Costs

As of 1990 carbon-13 was commercially available at a cost of about $700 / g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Nitrogen

Nitrogen compounds for rocket fuels.

Atomic Number:	7
Atomic Symbol:	N
Atomic Weight:	14.00674
Electron Configuration:	$[He]2s^2 2p^3$

History

(L. *nitrum*, Gr. *Nitron*, native soda; genes, forming) Nitrogen was discovered by chemist and physician Daniel Rutherford in 1772. He removed oxygen and carbon dioxide from air and showed that the residual gas would not support combustion or living organisms. At the same time there were other noted scientists working on the problem of nitrogen. These included Scheele, Cavendish, Priestley, and others. They called it “burnt or dephlogisticated air,” which meant air without oxygen.

Sources

Nitrogen gas (N_2) makes up 78.1% of the Earth’s air, by volume. The atmosphere of Mars, by comparison, is only 2.6% nitrogen. From an exhaustible source in our atmosphere, nitrogen gas can be obtained by liquefaction and fractional distillation. Nitrogen is found in all living systems as part of the makeup of biological compounds.

The Element

The French chemist Antoine Laurent Lavoisier named nitrogen *azote*, meaning *without life*. However, nitrogen compounds are found in foods, fertilizers, poisons, and explosives. Nitrogen, as a gas is colorless, odorless, and generally considered an inert element. As a liquid (boiling point = minus 195.8^{o}C), it is also colorless and odorless, and is similar in appearance to water. Nitrogen gas can be prepared by heating a water solution of ammonium nitrite (NH_4NO_3).

Nitrogen Compounds and Nitrogen in Nature

Sodium nitrate ($NaNO_3$) and potassium nitrate (KNO_3) are formed by the decomposition of organic matter with compounds of these metals present. In certain dry areas of the world these saltpeters are found in quantity and are used as fertilizers. Other inorganic nitrogen compounds are nitric acid (HNO_3), ammonia (NH_3), the oxides (NO, NO_2, N_2O_4, N_2O), cyanides (CN^-), etc.

The nitrogen cycle is one of the most important processes in nature for living organisms. Although nitrogen gas is relatively inert, bacteria in the soil are capable of “fixing” the nitrogen into a usable form (as a fertilizer) for plants. In other words, Nature has provided a method to produce nitrogen for plants to grow. Animals eat the plant material where the nitrogen has been incorporated into their system, primarily as protein. The cycle is completed when other bacterial convert the waste nitrogen compounds back to nitrogen gas. Nitrogen has become crucial to life being a component of all proteins.

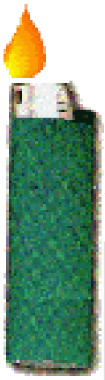
Ammonia

Ammonia (NH_3) is the most important commercial compound of nitrogen. It is produced by the Haber Process. Natural gas (methane, CH_4) is reacted with steam to produce carbon dioxide and hydrogen gas (H_2) in a two step process. Hydrogen gas and nitrogen gas are then reacted in the Haber Process to produce ammonia. This colorless gas with a pungent odor is easily liquefied. In fact, the liquid is used as a nitrogen fertilizer. Ammonia is also used in the production of urea, NH_2CONH_2, which is used as a fertilizer, in the plastic industry, and in the livestock industry as a feed supplement. Ammonia is often the starting compound for many other nitrogen compounds.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Oxygen

For combustion

Atomic Number:	8
Atomic Symbol:	O
Atomic Weight:	15.9994
Electron Configuration:	$[He]2s^2 2p^4$

History

(Gr. oxys, sharp, acid, and genes, forming; acid former) For many centuries, workers occasionally realized air was composed of more than one component. The behavior of oxygen and nitrogen as components of air led to the advancement of the phlogiston theory of combustion, which captured the minds of chemists for a century. Oxygen was prepared by several workers, including Bayen and Borch, but they did not know how to collect it, did not study its properties, and did not recognize it as an elementary substance.

Priestley is generally credited with its discovery, although Scheele also discovered it independently.

Its atomic weight was used as a standard of comparison for each of the other elements until 1961 when the International Union of Pure and Applied Chemistry adopted carbon 12 as the new basis.

Sources

Oxygen is the third most abundant element found in the sun, and it plays a part in the carbon-nitrogen cycle, the process once thought to give the sun and stars their energy. Oxygen under excited conditions is responsible for the bright red and yellow-green colors of the Aurora.

A gaseous element, oxygen forms 21% of the atmosphere by volume and is obtained by liquefaction and fractional distillation. The atmosphere of Mars contains about 0.15% oxygen. The element and its compounds make up 49.2%, by weight, of the earth's crust. About two thirds of the human body and nine tenths of water is oxygen.

In the laboratory it can be prepared by the electrolysis of water or by heating potassium chlorate with manganese dioxide as a catalyst.

Properties

The gas is colorless, odorless, and tasteless. The liquid and solid forms are a pale blue color and are strongly paramagnetic.

Forms

Ozone (O_3), a highly active compound, is formed by the action of an electrical discharge or ultraviolet light on oxygen.

Ozone's presence in the atmosphere (amounting to the equivalent of a layer 3 mm thick under ordinary pressures and temperatures) helps prevent harmful ultraviolet rays of the sun from reaching the earth's surface. Pollutants in the atmosphere may have a detrimental effect on this ozone layer. Ozone is toxic and exposure should not exceed 0.2 mg/m# (8-hour time-weighted average - 40-hour work week). Undiluted ozone has a bluish color. Liquid ozone is bluish black and solid ozone is violet-black.

Compounds

Oxygen, which is very reactive, is a component of hundreds of thousands of organic compounds and combines with most elements.

Uses

Plants and animals rely on oxygen for respiration. Hospitals frequently prescribe oxygen for patients with respiratory ailments.

Isotopes

Oxygen has nine isotopes. Natural oxygen is a mixture of three isotopes.

Natural occurring oxygen 18 is stable and available commercially, as is water (H2O with 15% 18O). Commercial oxygen consumption in the U.S. is estimated at 20 million short tons per year and the demand is expected to increase substantially.

Oxygen enrichment of steel blast furnaces accounts for the greatest use of the gas. Large quantities are also used in making synthesis gas for ammonia and methanol, ethylene oxide, and for oxy-acetylene welding.

Air separation plants produce about 99% of the gas, while electrolysis plants produce about 1%.

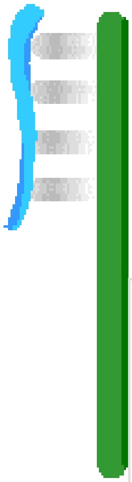
Costs

The gas costs 5 cents / ft^3 in small quantities, and about $15/ton in large quantities.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Fluorine

For toothpaste

Atomic Number:	9
Atomic Symbol:	F
Atomic Weight:	18.998403
Electron Configuration:	$[He]2s^2 2p^5$

History

(L. and F. fluere, flow or flux) In 1529, Georigius Agricola described the use of fluorspar as a flux, and as early as 1670 Schwandhard found that glass was etched when exposed to fluorspar treated with acid. Scheele and many later investigators, including Davy, Gay-Lussac, Lavoisier, and Thenard, experimented with hydrofluoric acid, some experiments ending in tragedy.

The element was finally isolated in 1866 by Moissan after nearly 74 years of continuous effort.

Properties

Fluorine is the most electronegative and reactive of all elements. It is a pale yellow, corrosive gas, which reacts with most organic and inorganic substances. Finely divided metals, glass, ceramics, carbon, and even water burn in fluorine with a bright flame.

Until World War II, there was no commercial production of elemental fluorine. The nuclear bomb project and nuclear energy applications, however, made it necessary to produce large quantities.

Uses

Fluorine and its compounds are used in producing uranium (from the hexafluoride) and more than 100 commercial fluorochemicals, including many well known high-temperature plastics. Hydrofluoric acid etches the glass of light bulbs, etc. Fluorochlorohydrocarbons are extensively used in air conditioning and refrigeration.

The presence of fluorine as a soluble fluoride in drinking water to the extent of 2 ppm may cause mottled enamel in teeth, when used by children acquiring permanent teeth; in smaller amounts, however, fluorides are added to water supplies to prevent dental cavities.

Elemental fluorine has been studied as a rocket propellant as it has an exceptionally high specific impulse value.

Compounds

One hypothesis says that fluorine can be substituted for hydrogen wherever it occurs in organic compounds, which could lead to an astronomical number of new fluorine compounds. Compounds of fluorine with rare gases have now been confirmed in fluorides of xenon, radon, and krypton.

Handling

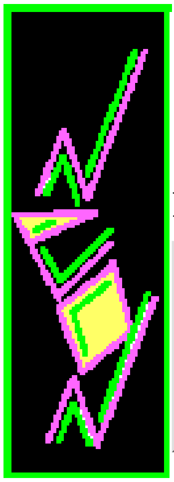
Elemental fluorine and the fluoride ion are highly toxic. The free element has a characteristic pungent odor, detectable in concentrations as low as 20 ppb, which is below the safe working level. The recommended maximum allowable concentration for a daily 8-hour time-weighted exposure is 1 ppm.

Safe handling techniques enable the transport liquid fluorine by the ton.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Neon

For lights

Atomic Number:	10
Atomic Symbol:	Ne
Atomic Weight:	20.179
Electron Configuration:	$[He]2s^2 2p^6$

History

(Gr. neos, new) Discovered by Ramsay and Travers in 1898. Neon is a rare gaseous element present in the atmosphere to the extent of 1 part in 65,000 of air. It is obtained by liquefaction of air and separated from the other gases by fractional distillation.

Isotopes

Natural neon is a mixture of three isotopes. Six other unstable isotopes are known.

Compounds

Neon, a very inert element, is however said to form a compound with fluorine. It is still questionable if true compounds of neon exist, but evidence is mounting in favor of their existence. The ions, Ne+, (NeAr)+, (NeH)+, and (HeNe+) are known from optical and mass spectrometric studies. Neon also forms an unstable hydrate.

Properties

In a vacuum discharge tube, neon glows reddish orange.

It has over 40 times more refrigerating capacity per unit volume than liquid helium and more than three times that of liquid hydrogen. It is compact, inert, and is less expensive than helium when it meets refrigeration requirements.

Of all the rare gases, the discharge of neon is the most intense at ordinary voltages and currents.

Uses

Although neon advertising signs account for the bulk of its use, neon also functions in high-voltage indicators, lightning arrestors, wave meter tubes, and TV tubes. Neon and helium are used in making gas lasers. Liquid neon is now commercially available and is finding important application as an economical cryogenic refrigerant.

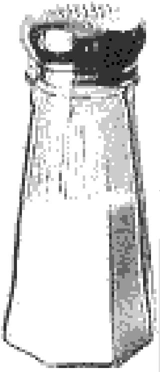
Costs

Neon costs about $2.00/l.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Sodium

For salt (sodium chloride)

Atomic Number:	11
Atomic Symbol:	Na
Atomic Weight:	22.98977
Electron Configuration:	$[Ne]3s^{1}$

History

(English, soda; Medieval Latin, sodanum, headache remedy) Long recognized in compounds, sodium was first isolated by Davy in 1807 by electrolysis of caustic soda.

Sources

Sodium is present in fair abundance in the sun and stars. The D lines of sodium are among the most prominent in the solar spectrum. Sodium is the fourth most abundant element on earth, comprising about 2.6% of the earth's crust; it is the most abundant of the alkali group of metals.

It is now obtained commercially by the electrolysis of absolutely dry fused sodium chloride. This method is much cheaper than that of electrolyzing sodium hydroxide, as was used several years ago.

Compounds

The most common compound is sodium chloride, but it occurs in many other minerals, such as soda niter, cryolite, amphibole, zeolite, etc.

Properties

Sodium, like every reactive element, is never found free in nature. Sodium is a soft, bright, silvery metal which floats on water, decomposing it with the evolution of hydrogen and the formation of the hydroxide. It may or may not ignite spontaneously on water, depending on the amount of oxide and metal exposed to the water. It normally does not ignite in air at temperatures below $115_{o}C$.

Uses

Metallic sodium is vital in the manufacture of esters and in the preparation of organic compounds. The metal may be used to improve the structure of certain alloys, to descale metal, and to purify molten metals.

An alloy of sodium with potassium, NaK, is also an important heat transfer agent.

Compounds

Sodium compounds are important to the paper, glass, soap, textile, petroleum, chemical, and metal industries. Soap is generally a sodium salt of certain fatty acids. The importance of common salt to animal nutrition has been recognized since prehistoric times.

Among the many compounds that are of the greatest industrial importance are common salt (NaCl), soda ash (Na_2CO_3), baking soda ($NaHCO_3$), caustic soda (NaOH), Chile saltpeter ($NaNO_3$), di- and tri-sodium phosphates, sodium thiosulfate (hypo, $Na_2S_2O_3 . 5H_2O$), and borax ($Na_2B_4O_7 . 10H_2O$).

Isotopes

Thirteen isotopes of sodium are recognized.

Cost

Metallic sodium is priced at about 15 to 20 cents/lb in quantity. Reagent grade (ACS) sodium in January 1990 cost about $35/lb. On a volume basis, it is the cheapest of all metals.

Handling

Sodium metal should be handled with great care. It cannot be maintained in an inert atmosphere and contact with water and other substances with which sodium reacts should be avoided.

Isotopes available at Los Alamos National Laboratory

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Magnesium

For racing bikes

Atomic Number:	12
Atomic Symbol:	Mg
Atomic Weight:	24.305
Electron Configuration:	$[Ne]3s^2$

History

(Magnesia, district in Thessaly) Compounds of magnesium have long been known. Black recognized magnesium as an element in 1755. Davy isolated it in 1808 and Bussy prepared it in coherent form in 1831. Magnesium is the eighth most abundant element in the earth's crust. It does not occur uncombined, but is found in large deposits in the form of magnesite, dolomite, and other minerals.

Sources

The metal is now principally obtained in the U.S. by electrolysis of fused magnesium chloride derived from brines, wells, and sea water.

Properties

Magnesium is a light, silvery-white, and fairly tough metal. It tarnishes slightly in air, and finely divided magnesium readily ignites upon heating in air and burns with a dazzling white flame.

Uses

Uses include flashlight photography, flares, and pyrotechnics, including incendiary bombs. It is one third lighter than aluminum, and in alloys is essential for airplane and missile construction. The metal improves the mechanical, fabrication, and welding characteristics of aluminum when used as an alloying agent. Magnesium is used in producing nodular graphite in cast iron, and is used as an additive to conventional propellants.

It is also used as a reducing agent in the production of pure uranium and other metals from their salts. The hydroxide (milk of magnesia), chloride, sulfate (Epsom salts), and citrate are used in medicine. Dead-burned magnesite is employed for refractory purposes such as brick and liners in furnaces and converters.

Compounds

Organic magnesium is important in both plant and animal life. Chlorophylls are magnesium-centered perphyrins.

The adult daily nutritional requirement, which is affected by various factors include weight and size, is about 300 mg / day.

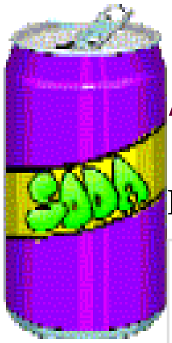
Handling

Because serious fires can occur, great care should be taken in handling magnesium metal, especially in the finely divided state. Water should not be used on burning magnesium or on magnesium fires.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Aluminum

For soda cans.

Atomic Number:	13
Atomic Symbol:	Al
Atomic Weight:	26.98154
Electron Configuration:	2-8-3

History

(L. alumen, alum) The ancient Greeks and Romans used alum as an astringent and as a mordant in dyeing. In 1761 de Morveau proposed the name alumine for the base in alum, and Lavoisier, in 1787, thought this to be the oxide of a still undiscovered metal.

Wohler is generally credited with having isolated the metal in 1827, although an impure form was prepared by Oersted two years earlier. In 1807, Davy proposed the name aluminum for the metal, undiscovered at that time, and later agreed to change it to aluminum. Shortly thereafter, the name aluminum was adopted to conform with the "ium" ending of most elements, and this spelling is now in use elsewhere in the world.

Aluminium was also the accepted spelling in the U.S. until 1925, at which time the American Chemical Society officially decided to use the name aluminum thereafter in their publications.

Sources

The method of obtaining aluminum metal by the electrolysis of alumina dissolved in cryolite was discovered in 1886 by Hall in the U.S. and at about the same time by Heroult in France. Cryolite, a natural ore found in Greenland, is no longer widely used in commercial production, but has been replaced by an artificial mixture of sodium, aluminum, and calcium fluorides.

Aluminum can now be produced from clay, but the process is not economically feasible at present. Aluminum is the most abundant metal to be found in the earth's crust (8.1%), but is never found free in nature. In addition to the minerals mentioned above, it is found in granite and in many other common minerals.

Properties

Pure aluminum, a silvery-white metal, possesses many desirable characteristics. It is light, it is nonmagnetic and nonsparking, stands second among metals in the scale of malleability, and

sixth in ductility.

Uses

It is extensively used for kitchen utensils, outside building decoration, and in thousands of industrial applications where a strong, light, easily constructed material is needed.

Although its electrical conductivity is only about 60% that of copper, it is used in electrical transmission lines because of its light weight. Pure aluminum is soft and lacks strength, but it can be alloyed with small amounts of copper, magnesium, silicon, manganese, and other elements to impart a variety of useful properties.

These alloys are of vital importance in the construction of modern aircraft and rockets. Aluminum, evaporated in a vacuum, forms a highly reflective coating for both visible light and radiant heat. These coatings soon form a thin layer of the protective oxide and do not deteriorate as do silver coatings. They are used to coat telescope mirrors and to make decorative paper, packages, toys.

Compounds

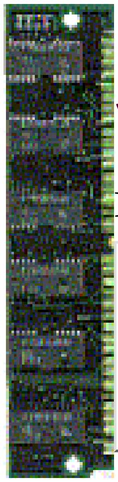
The compounds of greatest importance are aluminum oxide, the sulfate, and the soluble sulfate with potassium (alum). The oxide, alumina, occurs naturally as ruby, sapphire, corundum, and emery, and is used in glassmaking and refractories. Synthetic ruby and sapphire are used in lasers for producing coherent light.

Isotopes

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Silicon

For micro chips.

Atomic Number:	14
Atomic Symbol:	Si
Atomic Weight:	28.086
Electron Configuration:	$[Ne]3s^2 3p^2$

History

(L. silex, silicis, flint) Davy in 1800 thought silica to be a compound and not an element; later in 1811, Gay Lussac and Thenard probably prepared impure amorphous silicon by heating potassium with silicon tetrafluoride.

In 1824 Berzelius, generally credited with the discovery, prepared amorphous silicon by the same general method and purified the product by removing the fluosilicates by repeated washings. Deville in 1854 first prepared crystalline silicon, the second allotropic form of the element.

Sources

Silicon is present in the sun and stars and is a principal component of a class of meteorites known as *aerolites*. It is also a component of tektites, a natural glass of uncertain origin.

Silicon makes up 25.7% of the earth's crust, by weight, and is the second most abundant element, being exceeded only by oxygen. Silicon is not found free in nature, but occurs chiefly as the oxide and as silicates. Sand, quartz, rock crystal, amethyst, agate, flint, jasper, and opal are some of the forms in which the oxide appears. Granite, hornblende, asbestos, feldspar, clay, mica, etc. are but a few of the numerous silicate minerals.

Silicon is prepared commercially by heating silica and carbon in an electric furnace, using carbon electrodes. Several other methods can be used for preparing the element. Amorphous silicon can be prepared as a brown powder, which can be easily melted or vaporized. The Czochralski process is commonly used to produce single crystals of silicon used for solid-state or semiconductor devices. Hyperpure silicon can be prepared by the thermal decomposition of ultra-pure trichlorosilane in a hydrogen atmosphere, and by a vacuum float zone process.

Uses

Silicon is one of man's most useful elements. In the form of sand and clay it is used to make concrete and brick; it is a useful refractory material for high-temperature work, and in the form of silicates it is used in making enamels, pottery, etc. Silica, as sand, is a principal ingredient of glass, one of the most inexpensive of materials with excellent mechanical, optical, thermal, and electrical properties. Glass can be made in a very great variety of shapes, and is used as containers, window glass, insulators, and thousands of other uses. Silicon tetrachloride can be used as iridize glass.

Hyperpure silicon can be doped with boron, gallium, phosphorus, or arsenic to produce silicon for use in transistors, solar cells, rectifiers, and other solid-state devices which are used extensively in the electronics and space-age industries.

Hydrogenated amorphous silicon has shown promise in producing economical cells for converting solar energy into electricity.

Silicon is important to plant and animal life. Diatoms in both fresh and salt water extract Silica from the water to build their cell walls. Silica is present in the ashes of plants and in the human skeleton. Silicon is an important ingredient in steel; silicon carbide is one of the most important abrasives and has been used in lasers to produce coherent light of 4560 A.

Silcones are important products of silicon. They may be prepared by hydrolyzing a silicon organic chloride, such as dimethyl silicon chloride. Hydrolysis and condensation of various substituted chlorosilanes can be used to produce a very great number of polymeric products, or silicones, ranging from liquids to hard, glasslike solids with many useful properties.

Properties

Crystalline silicon has a metallic luster and grayish color. Silicon is a relatively inert element, but it is attacked by halogens and dilute alkali. Most acids, except hydrofluoric, do not affect it. Elemental silicon transmits more than 95% of all wavelengths of infrared, from 1.3 to 6.y micro-m.

Costs

Regular grade silicon (99%) costs about $0.50 / g. Silicon 99.9% pure costs about $50 / lb; hyperpure silicon may cost as much as $100 / oz.

Handling

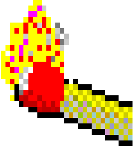
Miners, stonecutters, and others engaged in work where siliceous dust is breathed into large quantities often develop a serious lung disease known as silicosis.

Isotopes

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Phosphorus

For matches

Atomic Number:	15
Atomic Symbol:	P
Atomic Weight:	30.97376
Electron Configuration:	$[Ne]3s^2 3p^3$

History

(Gr. phosphoros, light bearing; ancient name for the planet Venus when appearing before sunrise) Brand discovered phosphorus in 1669 by preparing it from urine.

Properties

Phosphorus exists in four or more allotropic forms: white (or yellow), red, and black (or violet). Ordinary phosphorus is a waxy white solid; when pure it is colorless and transparent. White phosphorus has two modifications: alpha and beta with a transition temperature at $-3.8^{o}C$.

It is insoluble in water, but soluble in carbon disulfide. It takes fire spontaneously in air, burning to the pentoxide.

Sources

Never found free in nature, it is widely distributed in combination with minerals. Phosphate rock, which contains the mineral apatite, an impure tri-calcium phosphate, is an important source of the element. Large deposits are found in Russia, in Morocco, and in Florida, Tennessee, Utah, Idaho, and elsewhere.

Handling

It is very poisonous, 50 mg constituting an approximate fatal dose. Exposure to white phosphorus should not exceed 0.1 mg/m^3 (8-hour time-weighted average - 40-hour work week). White phosphorus should be kept under water, as it is dangerously reactive in air, and it should be handled with forceps, as contact with the skin may cause severe burns.

When exposed to sunlight or when heated in its own vapor to $250^{o}C$, it is converted to the red variety, which does not phosphoresce in air as does the white variety. This form does not ignite

spontaneously and is not as dangerous as white phosphorus. It should, however, be handled with care as it does convert to the white form at some temperatures and it emits highly toxic fumes of the oxides of phosphorus when heated. The red modification is fairly stable, sublimes with a vapor pressure of 1 atm at 17C, and is used in the manufacture of safety matches, pyrotechnics, pesticides, incendiary shells, smoke bombs, tracer bullets, etc.

Production

White phosphorus may be made by several methods. By one process, tri-calcium phosphate, the essential ingredient of phosphate rock, is heated in the presence of carbon and silica in an electric furnace or fuel-fired furnace. Elementary phosphorus is liberated as vapor and may be collected under phosphoric acid, an important compound in making super-phosphate fertilizers.

Uses

In recent years, concentrated phosphoric acids, which may contain as much as 70% to 75% P_2O_5 content, have become of great importance to agriculture and farm production. World-wide demand for fertilizers has caused record phosphate production. Phosphates are used in the production of special glasses, such as those used for sodium lamps.

Bone-ash, calcium phosphate, is used to create fine chinaware and to produce mono-calcium phosphate, used in baking powder.

Phosphorus is also important in the production of steels, phosphor bronze, and many other products. Trisodium phosphate is important as a cleaning agent, as a water softener, and for preventing boiler scale and corrosion of pipes and boiler tubes.

Phosphorus is also an essential ingredient of all cell protoplasm, nervous tissue, and bones.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Sulfur

For fire works.

Atomic Number:	16
Atomic Symbol:	S
Atomic Weight:	32.06
Electron Configuration:	$[Ne]3s^2 3p^4$

History

(Sanskrit, sulvere; L. sulpur) Known to the ancients; referred to in Genesis as brimstone.

Sources

Sulfur is found in meteorites. R.W. Wood suggests that the dark area near the crater Aristarchus is a sulfur deposit.

Sulfur occurs native in the vicinity of volcanos and hot springs. It is widely distributed in nature as iron pyrites, galena, sphalerite, cinnabar, stibnite, gypsum, epsom salts, celestite, barite, etc.

Production

Sulfur is commercially recovered from wells sunk into the salt domes along the Gulf Coast of the U.S. Using the Frasch process heated water is forced into the wells to melt the sulfur, which is then brought to the surface.

Sulfur also occurs in natural gas and petroleum crudes and must be removed from these products. Formerly this was done chemically, which wasted the sulfur; new processes now permit recovery. Large amounts of sulfur are being recovered from Alberta gas fields.

Properties

Sulfur is pale yellow, odorless, brittle solid, which is insoluble in water but soluble in carbon disulfide. In every state, whether gas, liquid or solid, elemental sulfur occurs in more than one allotropic form or modification; these present a confusing multitude of forms whose relations are not yet fully understood.

In 1975, University of Pennsylvania scientists reported synthesis of polymeric sulfur nitride,

which has the properties of a metal, although it contains no metal atoms. The material has unusual optical and electrical properties.

High-purity sulfur is commercially available in purities of 99.999+%.

Amorphous or "plastic" sulfur is obtained by fast cooling of the crystalline form. X-ray studies indicate that amorphous sulfur may have a helical structure with eight atoms per spiral. Crystalline sulfur seems to be made of rings, each containing eight sulfur atoms, which fit together to give a normal X-ray pattern.

Isotopes

Eleven isotopes of sulfur exist. None of the four isotopes that in nature are radioactive. A finely divided form of sulfur, known as flowers of sulfur, is obtained by sublimation.

Compounds

Organic compounds containing sulfur are very important. Calcium sulfur, ammonium sulfate, carbon disulfide, sulfur dioxide, and hydrogen sulfide are but a few of the many important compounds of sulfur.

Uses

Sulfur is a component of black gunpowder, and is used in the vulcanization of natural rubber and a fungicide. It is also used extensively in making phosphatic fertilizers. A tremendous tonnage is used to produce sulfuric acid, the most important manufactured chemical.

It is used to make sulfite paper and other papers, to fumigate fumigant, and to bleach dried fruits. The element is a good insulator.

Sulfur is essential to life. It is a minor constituent of fats, body fluids, and skeletal minerals.

Handling

Carbon disulfide, hydrogen sulfide, and sulfur dioxide should be handled carefully. Hydrogen sulfide in small concentrations can be metabolized, but in higher concentrations it quickly can cause death by respiratory paralysis.

It quickly deadens the sense of smell. Sulfur dioxide is a dangerous component in atmospheric air pollution.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Chlorine

For water purification.

Atomic Number:	17
Atomic Symbol:	Cl
Atomic Weight:	35.453
Electron Configuration:	$[Ne]3s^2 3p^5$

History

(Gr. chloros, greenish yellow) Discovered in 1774 by Scheele, who thought it contained oxygen. Chlorine was named in 1810 by Davy, who insisted it was an element.

Sources

In nature it is found in the combined state only, chiefly with sodium as common salt (NaCl), carnallite, and sylvite.

Properties

It is a member of the halogen (salt-forming) group of elements and is obtained from chlorides by the action of oxidizing agents and more often by electrolysis; it is a greenish-yellow gas, combining directly with nearly all elements. At $10^{\circ}C$ one volume of water dissolves 3.10 volumes of chlorine, at $30^{\circ}C$ only 1.77 volumes.

Uses

Chlorine is widely used in making many everyday products. It is used for producing safe drinking water the world over. Even the smallest water supplies are now usually chlorinated.

It is also extensively used in the production of paper products, dyestuffs, textiles, petroleum products, medicines, antiseptics, insecticides, food, solvents, paints, plastics, and many other consumer products.

Most of the chlorine produced is used in the manufacture of chlorinated compounds for sanitation, pulp bleaching, disinfectants, and textile processing. Further use is in the manufacture of chlorates, chloroform, carbon tetrachloride, and in the extraction of bromine.

Organic chemistry demands much from chlorine, both as an oxidizing agent and in substitution, since it often brings many desired properties in an organic compound when substituted for hydrogen, as in one form of synthetic rubber.

Handling

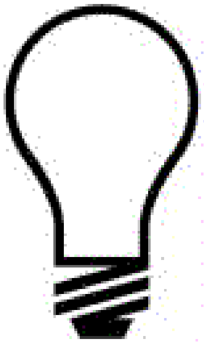
Chlorine is a respiratory irritant. The gas irritates the mucus membranes and the liquid burns the skin. As little as 3.5 ppm can be detected as an odor, and 1000 ppm is likely to be fatal after a few deep breaths. In fact, chlorine was used as a war gas in 1915.

Exposure to chlorine should not exceed 0.5 ppm (8-hour time-weighted average - 40 hour week.)

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Argon

For light bulbs.

Atomic Number:	18
Atomic Symbol:	Ar
Atomic Weight:	39.948
Electron Configuration:	[Ne]$3s^2 3p^6$

History

(Gr. argos, inactive) Its presence in air was suspected by Cavendish in 1785, discovered by Lord Rayleigh and Sir William Ramsay in 1894.

Sources

The gas is prepared by fractionation of liquid air because the atmosphere contains 0.94% argon. The atmosphere of Mars contains 1.6% of 40Ar and 5 p.p.m. of 36Ar.

Properties

Argon is two and one half times as soluble in water as nitrogen, having about the same solubility as oxygen. Argon is colorless and odorless, both as a gas and liquid. Argon is considered to be a very inert gas and is not known to form true chemical compounds, as do krypton, xenon, and radon.

Isotopes

Naturally occurring argon is a mixture of three isotopes. Twelve other radioactive isotopes are known to exist.

Uses

It is used in electric light bulbs and in fluorescent tubes at a pressure of about 400 Pa. and in filling photo tubes, glow tubes, etc. Argon is also used as an inert gas shield for arc welding and cutting, as blanket for the production of titanium and other reactive elements, and as a protective atmosphere for growing silicon and germanium crystals.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Potassium

For fertilizer.

Atomic Number:	19
Atomic Symbol:	K
Atomic Weight:	39.098
Electron Configuration:	$[Ar]4s^1$

History

(English, potash - pot ashes; L.. kalium, Arab qali, alkali) Discovered in 1807 by Davy, who obtained it from caustic potash (KOH); this was the first metal isolated by electrolysis.

Sources

The metal is the seventh most abundant and makes up about 2.4% by weight of the earth's crust. Most potassium minerals are insoluble and the metal is obtained from them only with great difficulty.

Certain minerals, however, such as sylvite, carnallite, langbeinite, and polyhalite are found in ancientlake and sea beds and form rather extensive deposits from which potassium and its salts can readily be obtained. Potash is mined in Germany, New Mexico, California, Utah, and elsewhere. Large deposits of potash, found at a depth of some 3000 ft in Saskatchewan, promise to be important in coming years.

Potassium is also found in the ocean, but is present only in relatively small amounts, compared to sodium.

Production

Potassium is never found free in nature, but is obtained by electrolysis of the hydroxide, much in the same manner as prepared by Davy. Thermal methods also are commonly used to produce potassium (such as by reduction of potassium compounds with CaC_2, C, Si, or Na).

Uses

The greatest demand for potash has been in its use for fertilizers. Potassium is an essential constituent for plant growth and is found in most soils.

An alloy of sodium and potassium (NaK) is used as a heat-transfer medium. Many potassium salts are of utmost importance, including the hydroxide, nitrate, carbonate, chloride, chlorate, bromide, iodide, cyanide, sulfate, chromate, and dichromate.

Properties

It is one of the most reactive and electropositive of metals. Except for lithium, it is the lightest known metal. It is soft, easily cut with a knife, and is silvery in appearance immediately after a fresh surface is exposed. It rapidly oxidizes in air and must be preserved in a mineral oil such as kerosene.

As with other metals of the alkali group, it decomposes in water with the evolution of hydrogen. It catches fire spontaneously on water. Potassium and its salts impart a violet color to flames.

Isotopes

Seventeen isotopes of potassium are known. Ordinary potassium is composed of three isotopes, one of which is ^{40}oK (0.0118%), a radioactive isotope with a half-life of 1.28 x 10^{9} years.

Handling

The radioactivity presents no appreciable hazard.

Cost

Metallic potassium is available commercially for about $40/lb in small quantities.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Calcium

For cement and plaster of paris.

Atomic Number:	20
Atomic Symbol:	Ca
Atomic Weight:	40.08
Electron Configuration:	[Ar]$4s^2$

History

(L. calx, lime) Though lime was prepared by the Romans in the first century under the name calx, the metal was not discovered until 1808. After learning that Berzelius and Pontin prepared calcium amalgam by electrolyzing lime in mercury, Davy was able to isolate the impure metal.

Sources

Calcium is a metallic element, fifth in abundance in the earth's crust, of which if forms more than 3%. It is an essential constituent of leaves, bones, teeth, and shells. Never found in nature uncombined, it occurs abundantly as limestone, gypsum, and fluorite. Apatite is the fluorophosphate or chlorophosphate of calcium.

Properties

The metal has a silvery color, is rather hard, and is prepared by electrolysis of the fused chloride to which calcium fluoride is added to lower the melting point.

Chemically it is one of the alkaline earth elements; it readily forms a white coating of nitride in air, reacts with water, burns with a yellow-red flame, forming largely the nitride.

Uses

The metal is used as a reducing agent in preparing other metals such as thorium, uranium, zirconium, etc., and is used as a deoxidizer, desulfurizer, or decarburizer for various ferrous and nonferrous alloys. It is also used as an alloying agent for aluminum, beryllium, copper, lead, and magnesium alloys, and serves as a "getter" for residual gases in vacuum tubes, etc.

Compounds

Its natural and prepared compounds are widely used. Quicklime (CaO), which is made by heating limestone that is changed into slaked lime by carefully adding water, is the great base of chemical refinery with countless uses.

Mixed with sand it hardens as mortar and plaster by taking up carbon dioxide from the air. Calcium from limestone is an important element in Portland cement.

The solubility of the carbonate in water containing carbon dioxide causes the formation of caves with stalagtites and stalagmites and is responsible for hardness in water. Other important compounds are the carbide, chloride, cyanamide, hypochlorite, nitrate, and sulfide.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Scandium

For stadium lighting.

Atomic Number:	21
Atomic Symbol:	Sc
Atomic Weight:	44.9559
Electron Configuration:	$[Ar]4s^{2}3d^{1}$

History

(L. Scandia, Scandinavia) On the basis of the Periodic System, Mendeleev predicted the existence of ekaboron, which would have an atomic weight between 40 of calcium and 48 of titanium.

The element was discovered by Nilson in 1878 in the minerals euxenite and gadolinite, which had not yet been found anywhere except in Scandinavia. By processing 10 kg of euxenite and other residues of rare-earth minerals, Nilson was able to prepare about 2g of highly pure scandium oxide. Later scientists pointed out that Nilson's scandium was identical with Mendeleev's ekaboron.

Sources

Scandium is apparently much more abundant (the 23rd most) in the sun and certain stars than on earth (the 50th most abundant). It is widely distributed on earth, occurring in very minute quantities in over 800 mineral species. The blue color of beryl (aquamarine variety) is said to be due to scandium. It occurs as a principal component in the rare mineral thortveitite, found in Scandinavia and Malagasy. It is also found in the residues remaining after the extraction of tungsten from Zinnwald wolframite, and in wiikite and bazzite.

Most scandium is presently being recovered from thortveitite or is extracted as a by-product from uranium mill tailings. Metallic scandium was first prepared in 1937 by Fischer, Brunger, and Grienelaus who electrolyzed a eutectic melt of potassium, lithium, and scandium chlorides at 700 to 800°C. Tungsten wire and a pool of molten zinc served as the electrodes in a graphite crucible. Pure scandium is now produced by reducing scandium fluoride with calcium metal.

The production of the first pound of 99% pure scandium metal was announced in 1960.

Properties

Scandium is a silver-white metal which develops a slightly yellowish or pinkish cast upon exposure to air. A relatively soft element, scandium resembles yttrium and the rare-earth metals more than it resembles aluminum or titanium.

It is a very light metal and has a much higher melting point than aluminum, making it of interest to designers of spacecraft. Scandium is not attacked by a 1:1 mixture of HNO_3 and 48% HF.

Costs

Scandium oxide costs about $75/g.

Uses

About 20 kg of scandium (as Sc_2O_3) are now being used yearly in the U.S. to produce high-intensity lights, and the radioactive isotope 46Sc is used as a tracing agent in refinery crackers for crude oil, etc.

Scandium iodide added to mercury vapor lamps produces a highly efficient light source resembling sunlight, which is important for indoor or night-time color TV.

Handling

Little is yet known about the toxicity of scandium; therefore it should be handled with care.

Isotopes available at Los Alamos National Laboratory

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Tungsten

For lamp filaments.

Atomic Number:	74
Atomic Symbol:	W
Atomic Weight:	183.5
Electron Configuration:	[Xe]$6s^2 4f^{14} 5d^4$

History

(Swedish, tung sten, heavy stone) In 1779 Peter Woulfe examined the mineral now known as wolframite and concluded it must contain a new substance. Scheele, in 1781, found that a new acid could be made from tungsten (a name first applied about 1758 to a mineral now known as scheelite). Scheele and Berman suggested the possibility of obtaining a new metal by reducing this acid. The de Elhuyar brothers found acid in wolframite in 1783 that was identical to the acid of tungsten (tungstic acid) of Scheele, and in that year they succeeded in obtaining the element by reduction of this acid with charcoal. Tungsten occurs in wolframite, scheelite, huebnertie, and ferberite. Important deposits of tungsten occur in California, Colorado, South Korea, Bolivia, Russia, and Portugal. China is reported to have about 75% of the world's tungsten resources. Natural tungsten contains five stable isotopes. Twenty one other unstable isotopes are recognized. The metal is obtained commercially be recucing tungsten oxide with hydrogen or carbon. Pure tungsten is a stelel-gray to tin-white metal. Very pure tungsten can be cut with a hacksaw, and can be forged, spun, drawn, and extruded. The impure metal is brittle and can be worked only with difficulty. Tungsten has the highest melting point of all metals, and at temperatures over 1650$_o$C has the highest tensile strength. The metal oxidizes in air and must be protected at elevated temperatures. It has excellent corrosion resistance and is attacked only slightly by most mineral acids. The thermal expansion is about the same as borosilicate glass, which makes the metal useful for glass-to-metal seals. tungsten and its alloys are used extensively for filaments for electric lamps, electron and television tubes, and for metal evaporation work; for electrical contact points for automobile distributors; X-ray targets; windings and heating elements for electrical furnaces; and for numerous spacecraft and high-temperature applications. High-speed tool steels, Hastelloy(R), Stellite(R), and many other alloys contain tungsten. Tungsten carbide is of great importance to the metal-working, mining, and petroleum industries. Calcium and magnesium tungstates are widely used in fluorescent lighting; other salts of tungsten are used in the chemical and tanning industries. Tungsten disulfide is a dry, high-temperature lubricant, stable to 500C. Tungsten bronzes and other tungsten compounds are used in paints. Tungsten powder (99.9%) costs about $50/lb.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 11/8/96, CST Information Services Team

Periodic Table

Zinc

For gutters.

Atomic Number:	30
Atomic Symbol:	Zn
Atomic Weight:	65.38
Electron Configuration:	[Ar]$4s^2 3d^{10}$

History

(German Zink, of obscure origin) Centuries before zinc was recognized as a distinct element, zinc ores were used for making brass. Tubal-Cain, seven generations from Adam, is mentioned as being an "instructor in every artificer in brass and iron." An alloy containing 87 percent zinc has been found in prehistoric ruins in Transylvania.

Metallic zinc was produced in the 13th century A.D. in India by reducing calamine with organic substances such as wool. The metal was rediscovered in Europe by Marggraf in 1746, who showed that it could be obtained by reducing calamine with charcoal.

Sources

The principal ores of zinc are sphalerite (sulfide), smithsonite (carbonate), calamine (silicate), and franklinite (zine, manganese, iron oxide). One method of zinc extraction involves roasting its ores to form the oxide and reducing the oxide with coal or carbon, with subsequent distillation of the metal.

Isotopes

Naturally occurring zinc contains five stable isotopes. Sixteen other unstable isotopes are recognized.

Properties

Zinc is a bluish-white, lustrous metal. It is brittle at ordinary temperatures but malleable at 100 to 150$_{o}$C. It is a fair conductor of electricity, and burns in air at high red heat with evolution of white clouds of the oxide.

It exhibits superplasticity. Neither zinc nor zirconium is ferromagnetic; but $ZrZn_2$ exhibits

ferromagnetism at temperatures below 35$_{o}$K. It has unusual electrical, thermal, optical, and solid-state properties that have not been fully investigated.

Uses

The metal is employed to form numerous alloys with other metals. Brass, nickel silver, typewriter metal, commercial bronze, spring bronze, German silver, soft solder, and aluminum solder are some of the more important alloys.

Large quantities of zinc are used to produce die castings, which are used extensively by the automotive, electrical, and hardware industries. An alloy called Prestal(R), consisting of 78 percent zinc and 22 percent aluminum, is reported to be almost as strong as steel and as easy to mold as plastic. The alloy said to be so moldable that it can be molded into form using inexpensive ceramics or cement die casts.

Zinc is also used extensively to galvanize other metals such as iron to prevent corrosion. Zinc oxide is a unique and very useful material for modern civilization. It is widely used in the manufacture of paints, rubber products, cosmetics, pharmaceuticals, floor coverings, plastics, printing inks, soap, storage batteries, textiles, electrical equipment, and other products. Lithopone, a mixture of zinc sulfide and barium sulfate, is an important pigment.

Zinc sulfide is used in making luminous dials, X-ray and TV screens, and fluorescent lights.

The chloride and chromate are also important compounds. Zinc is an essential element in the growth of human beings and animals. Tests show that zinc-deficient animals require 50 percent more food to gain the same weight as an animal supplied with sufficient zinc.

Handling

Zinc is not considered to be toxic, but when freshly formed ZnO is inhaled a disorder known as the *oxide shakes* or *zinc chills* sometimes occurs. Where zinc oxide is encountered, recommendations include providing good ventilation to avoid concentration exceeding 5 mg/m$_3$, (time-weighted over an 8-hour exposure, 40-hour work week).

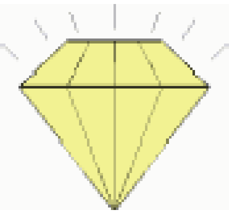
Costs

The price of zinc was roughly $0.70/lb in January 1990.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Zirconium

For zircon gemstone.

Atomic Number:	40
Atomic Symbol:	Zr
Atomic Weight:	91.22
Electron Configuration:	$[Kr]5s^2 4d^2$

History

(Persian zargun, gold like) The name zircon probably originated from the Persion word zargun, which describes the color of the gemstone now known as zircon, jargon, hyacinth, jacinth, or ligure. This mineral, or its variations, is mentioned in biblical writings. The mineral was not known to contain a new element until Klaproth, in 1789, analyzed a jargon from from Ceylon and found a new earth, which Werner named zircon (silex circonius), and Klaproth called Zirkonertz (zirconia). The impure metal was first isolated by Berzelius in 1824 by heating a mixture of potassium and potassium zirconium fluoride in a small decomposition process they developed.

Sources

Zirconium is found in abundance in S-type stars, and has been identified in the sun and meteorites. Analysis of lunar rock samples obtained during the various Apollo missions to the moon show a surprisingly high zirconium oxide content, compared with terrestrial rocks.

Isotopes

Naturally occurring zirconium contains five isotopes. Fifteen other isotopes are known to exist. Zircon, $ZrSiO_4$, the principal ore, is pure ZrO_2 in crystalline form having a hafnium content of about 1%. Zirconium also occurs in some 30 other recognized mineral species. Zirconium is produced commercially by reduction of chloride with magnesium (the Kroll Process), and by other methods. It is a grayish-white lustrous metal. When finely divided, the metal may ignite spontaneously in air, especially at elevated temperatures. The solid metal is much more difficult to ignite. The inherent toxicity of zirconium compounds is low. Hafnium is invariably found in zirconium ores, and the separation is difficult.

Commercial-grade zirconium contains from 1 to 3% hafnium. Zirconium has a low absorption cross section for neutrons, and is therefore used for nuclear energy applications, such as for

cladding fuel elements. Commercial nuclear power generation now takes more than 90% of zirconium metal production. Reactors of the commercial size, now being made, may use as much as a half-million linear feet of zirconium alloy tubing.

Properties

Reactor-grade zirconium is essentially free of hafnium. Zircaloy(R) is an important alloy developed specifically for nuclear applications. Zirconium is exceptionally resistant to corrosion by many common acids and alkalis, by sea water, and by other agents. Alloyed with zinc, zirconium becomes magnetic at temperatures below 35_oK.

Uses

It is used extensively by the chemical industry where corrosive agents are employed. Zirconium is used as a getter in vacuum tubes, as an alloying agent in steel, in surgical appliances, photoflash bulbs, explosive primers, rayon spinnerets, lamp filaments, etc. It is used in poison ivy lotions in the form of the carbonate as it combines with urushiol. With niobium, zirconium is superconductive at low temperatures and is used to make superconductive magnets, which offer hope of direct large-scale generation of electric power. Zirconium oxide (zircon) has a high index of refraction and is used as a gem material. The impure oxide, zirconia, is used for laboratory crucibles that will withstand heat shock, for linings of metallurgical furnaces, and by the glass and ceramic industries as a refractory material. Its use as a refractory material accounts for a large share of all zirconium consumed.

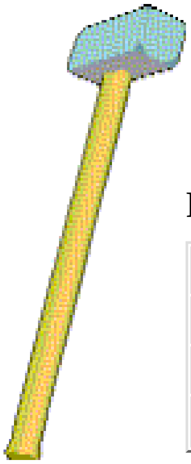
Cost

Zirconium of about 99.6% purity is available at a cost of about $150/kg.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Iron

For tools.

Atomic Number:	26
Atomic Symbol:	Fe
Atomic Weight:	55.847
Electron Configuration:	$[Ar]4s^2 3d^6$

History

(Anglo-Saxon, iron; L. ferrum) Iron was used prehistorically:

- Genesis mentions that Tubal-Cain, seven generations from Adam, was "an instructor of every artificer in brass and iron."
- A remarkable iron pillar, dating to about A.D. 400, remains standing today in Delhi, India. This solid shaft of wrought iron is about 7 1/4 m high by 40 cm in diameter. Corrosion to the pillar has been minimal although it has been exposed to the weather since its erection.

Sources

Iron is a relatively abundant element in the universe. It is found in the sun and many types of stars in considerable quantity. Its nuclei are very stable. Iron is a principal component of a meteorite class known as *siderites* and is a minor constituent of the other two meteorite classes. The core of the earth -- 2150 miles in radius -- is thought to be largely composed of iron with about 10 percent occluded hydrogen. The metal is the fourth most abundant element, by weight that makes up the crust of the earth.

The most common ore is hematite, which is frequently seen as black sands along beaches and banks of streams.

Isotopes

Common irons is a mixture of four isotopes. Ten other isotopes are known to exist.

Uses

Iron is a vital constituent of plant and animal life and appears in hemoglobin.

Taconite is becoming increasingly important as a commercial ore. The pure metal is not often encountered in commerce, but is usually alloyed with carbon or other metals.

Properties

The pure metal is very reactive chemically and rapidly corrodes, especially in moist air or at elevated temperatures. It has four allotropic forms or ferrites, known as alpha, beta, gamma, and omega, with transition points at 700, 928, and 1530C. The alpha form is magnetic, but when transformed into the beta form, the magnetism disappears although the lattice remains unchanged. The relations of these forms are peculiar. Pig iron is an alloy containing about 3 percent carbon with varying amounts of Sulfur, Silicon, Manganese, and Phosphorus.

Iron is hard, brittle, fairly fusible, and is used to produce other alloys, including steel. Wrought iron contains only a few tenths of a percent of carbon, is tough, malleable, less fusible, and has usually a "fibrous" structure.

Carbon steel is an alloy of iron with small amounts of Mn, S, P, and Si. Alloy steels are carbon steels with other additives such as nickel, chromium, vanadium, etc. Iron is a cheap, abundant, useful, and important metal.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Manganese

For plows.

Atomic Number:	25
Atomic Symbol:	Mn
Atomic Weight:	54.9380
Electron Configuration:	$[Ar]4s^2 3d^5$

History

(L. magnes, magnet, from magnetic properties of pyrolusite; It. manganese, corrupt form of magnesia)

Recognized by Scheele, Bergman, and others as an element and isolated by Gahn in 1774 by reduction of the dioxide with carbon.

Sources

Manganese minerals are widely distributed; oxides, silicates, and carbonates are the most common. The discovery of large quantities of manganese nodules on the floor of the oceans may become a source of manganese. These nodules contain about 24% manganese together with many other elements in lesser abundance.

Most manganese today is obtained from ores found in Russia, Brazil, Australia, Republic of S. Africa, Gabon, and India. Pyrolusite and rhodochrosite are among the most common manganese minerals. The metal is obtained by reduction of the oxide with sodium, magnesium, aluminum, or by elctrolysis.

Properties

It is gray-white, resembling iron, but is harder and very brittle. The metal is reactive chemically, and decomposes cold water slowly. Manganese is used to form many important alloys. In steel, manganese improves the rolling and forging qualities, strength, toughness, stiffness, wear resistance, hardness, and hardenability.

With aluminum and antimony, especially with small amounts of copper, it forms highly ferromagnetic alloys.

Manganese metal is ferromagnetic only after special treatment. The pure metal exists in four

allotropic forms. The alpha form is stable at ordinary temperature; gamma manganese, which changes to alpha at ordinary temperatures, is said to be flexible, soft, easily cut, and capable of being bent.

Uses

The dioxide (pyrolusite) is used as a depolarizer in dry cells, and is used to "decolorize" glass that is colored green by impurities of iron. Manganese by itself colors glass an amethyst color, and is responsible for the color of true amethyst. The dioxide is also used in the preparation of oxygen and chlorine, and in drying black paints. The permanganate is a powerful oxidizing agent and is used in quantitative analysis and in medicine.

Manganese is widely distributed throughout the animal kingdom. It is an important trace element and may be essential for utilization of vitamin B1.

Handling

Exposure to manganese dusts, fume, and compounds should not exceed the ceiling value of 5 mg/m^3 for even short periods because of the element's toxicity level.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Antimony

For ceramic glazes.

Atomic Number:	51
Atomic Symbol:	Sb
Atomic Weight:	121.75
Electron Configuration:	$[Kr]5s^{2}4d^{10}5p^{3}$

History

(Gr. anti plus monos - a metal not found alone) Antimony was recognized in compounds by the ancients and was known as a metal at the beginning of the 17th century and possibly much earlier.

Sources

It is not abundant, but is found in over 100 mineral species. It is sometimes found native, but more frequently as the sulfide stibnite.

Properties

It is a poor conductor of heat and electricity. Antimony and many of its compounds are toxic.

Uses

Antimony is finding use in semiconductor technology for making infrared detectors, diodes and Hall-effect devices. It greatly increases the hardness and mechanical strength of lead. Betteries, antifriction alloys, type metal, small arms and tracer bullets, cable sheathing, and minor products use about half the metal produced. Compounds taking up the other half are oxides, sulfides, sodium antimonate, and antimony trichloride. These are used in manufacturing flame-proofing compounds, paints ceramic enamels, glass, and pottery.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Copper

For pennies.

Atomic Number:	29
Atomic Symbol:	Cu
Atomic Weight:	63.546
Electron Configuration:	[Ar]$4s^1 3d^{10}$

History

(Latin *cuprum* , from the island of Cyprus) It is believed that copper has been mined for 5,000 years.

Properties

Copper is reddish and takes on a bright metallic luster. It is malleable, ductile, and a good conductor of heat and electricity (second only to silver in electrical conductivity).

Sources

Copper occasionally occurs native, and is found in many minerals such as cuprite, malachite, azurite, chalcopyrite, and bornite.

Large copper ore deposits are found in the U.S., Chile, Zambia, Zaire, Peru, and Canada. The most important copper ores are the sulfides, the oxides, and carbonates. From these, copper is obtained by smelting, leaching, and by electrolysis.

Uses

The electrical industry is one of the greatest users of copper. Iron's alloys -- brass and bronze -- are very important: all American coins are copper alloys and gun metals also contain copper.

Copper has wide use as an agricultural poison and as an algicide in water purification. Copper compounds, such as Fehling's solution, are widely used in analytical chemistry tests for sugar.

Availability

High-purity copper (99.999+ percent) is available commercially.

Isotopes available at Los Alamos National Laboratory

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Silver

For silverware.

Atomic Number:	47
Atomic Symbol:	Ag
Atomic Weight:	107.868
Electron Configuration:	$[Kr]5s^{1}4d^{10}$

History

(Anglo-Saxon, Seolfor siolfur; L. argentum) Silver has been known since ancient times. It is mentioned in Genesis. Slag dumps in Asia Minor and on islands in the Aegean Sea indicate that man learned to separate silver from lead as earl as 3000 B.C.

Sources

Silver occurs native and in ores such as argentite (Ag2S) and horn silver (AgCl); lead, lead-zinc, copper, gold, and copper-nickel ores are principal sources. Mexico, Canada, Peru, and the U.S. are the principal silver producers in the western hemisphere.

Production

Silver is also recovered during electrolytic refining of copper. Commercial fine silver contains at least 99.9% silver. Purities of 99.999+% are available commercially.

Properties

Pure silver has a brilliant white metallic luster. It is a little harder than gold and is very ductile and malleable, being exceeded only by gold and perhaps palladium. Pure silver has the highest electrical and thermal conductivity of all metals, and possesses the lowest contact resistance. It is stable in pure air and water, but tarnishes when exposed to ozone, hydrogen sulfide, or air containing sulfur. The alloys of silver are important.

Uses

Sterling silver is used for jewelry, silverware, etc. where appearance is paramount. This alloy contains 92.5% silver, the remainder being copper or some other metal. Silver is of the utmost importance in photography, about 30% of the U.S. industrial consumption going into this application. It is used for dental alloys. Silver is used in making solder and brazing alloys, electrical contacts, and high capacity silver-zinc and silver-cadmium batteries. Silver paints are used for making printed circuits. It is used in mirror production and may be deposited on glass or metals by chemical deposition, electrode position, or by evaporation. When freshly deposited, it is the best reflector of visible light known, but is rapidly tarnished and loses much of its reflectance. It is a poor reflector of ultraviolet. Silver fulminate, a powerful explosive, is sometimes formed during the silvering process. Silver iodide is used in seeding clouds to produce rain. Silver chloride has interesting optical properties as it can be made transparent; it also is a cement for glass. Silver nitrate, or lunar caustic, the most important silver compound, is used extensively in photography. Silver for centuries has been used traditionally for coinage by many countries of the world. In recent times, however, consumption of silver has greatly exceeded the output.

Handling

While silver itself is not considered to be toxic, most of its salts are poisonous. Exposure to silver (metal and soluble compounds, as Ag) in air should not exceed 0.01 mg/m^3, (8-hour time-weighted average - 40 hour week). Silver compounds can be absorbed in the circulatory system and reduced silver deposited in the various tissues of the body. A condition, known as argyria, results with a grayish pigmentation of the skin and mucous membranes. Silver has germicidal effects and kills many lower organisms effectively without harm to higher animals.

Cost

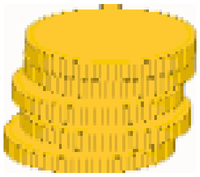
In 1939, the price of silver was fixed by the U.S. Treasury at 71 cents/troy oz., and at 90.5 cents/troy oz. in 1946. In November 1961 the U.S. Treasury suspended sales of nonmonetized silver, and the price stabilized for a time at about $1.29, the melt-down value of silver U.S. coins. The coinage act of 1965 authorized a change in the metallic composition of the three U.S. subsidiary denominations to clad or composite type coins. This was the first change in U.S. coinage since the monetary system was established in 1792. Clad dimes and quarters are made of an outer layer of 75% Cu and 25% Ni bonded to a central core of pure Cu. The composition of the one- and five-cent pieces remains unchanged. One-cent coins are 95% Cu and 5% Zn. Five-cent coins are 75% Cu and 25% Ni. Old silver dollars are 90% Ag and 10% Cu. Earlier subsidiary coins of 90% Ag and 10% Cu officially were to circulate alongside the old coins; however, in practice they have largely disappeared (Gresham's Law), as the value of the silver is now greater than their exchange value. Silver coins of other countries have largely been replaced with coins made of other metals. On June 24, 1968, the U.S. Government ceased to redeem U.S. Silver Certificates with silver. Since that time, the price of silver has fluctuated widely. As of January 1990, the price of silver was about $5.25/troy oz.; however, the price has fluctuated

considerably due to market instability.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Nickel

For coins.

Atomic Number:	28
Atomic Symbol:	Ni
Atomic Weight:	58.70
Electron Configuration:	$[Ar]4s^2 3d^8$

History

(German Nickel, Satan or Old Nick's and from kupfernickel, Old Nick's copper) Cronstedt discovered nickel in 1751 in kupfernickel (niccolite).

Sources

Nickel is found as a constitutent in most meteorites and often serves as one of the criteria for distinguishing a meteorite from other minerals. Iron meteorites, or siderites, may contain iron alloyed with from 5 percent to nearly 20 percent nickel. Nickel is obtained commercially from pentlandite and pyrrhotite of the Sudbury region of Ontario, a district that produces about 30 percent of the world's supply of nicke.

Other deposits are found in New Caledonia, Australia, Cuba, Indonesia, and elsewhere.

Properties

Nickel is silvery white and takes on a high polish. It is hard, malleable, ductile, somewhat ferromagnetic, and a fair conductor of heat and electricity. It belongs to the iron-cobalt group of metals and is chiefly valuable for the alloys it forms.

Uses

It is extensively used for making stainless steel and other corrosion-resistant alloys such as Invar(R), Monel(R), Inconel(R), and the Hastelloys(R). Tubing made of copper-nickel alloy is extensively used in making desalination plants for converting sea water into fresh water.

Nickel, used extensively to make coins and nickel steel for armor plates and burglar-proof vaults, and is also a component in Nichrome(R), Permalloy(R), and constantan.

Nickel gives glass a greenish color. Nickel plating is often used to provide a protective coating for other metals, and finely divided nickel is a catalyst for hydrogenating vegetable oils. It is also used in ceramics, in the manufacture of Alnico magnets, and in the Edison(R) storage battery.

Isotopes

The sulfate and the oxides are important compounds. Natural nickel is a mixture of five stable isotopes; nine other unstable isotopes are known.

Handling

Exposure to nickel metal and soluble compounds (as Ni) should not exceed 0.05 mg/cm^3 (8-hour time-weighted average - 40-hour work week). Nickel sulfide fume and dust is recognized as being potentially carcinogenic.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Chromium

For plating for cars.

Atomic Number:	24
Atomic Symbol:	Cr
Atomic Weight:	51.996
Electron Configuration:	$[Ar]4s^1 3d^5$

History

(Gr. chroma, color) Discovered in 1797 by Vauquelin, who prepared the metal the next year, chromium is a steel-gray, lustrous, hard metal that takes a high polish.

Sources

The principal ore is chromite, which is found in Zimbabwe, Russia, Transvaal, Turkey, Iran, Albania, Finland, Democratic Republic of Madagascar, and the Phillippines. The metal is usually produced by reducing the oxide with aluminum.

Uses

Chromium is used to harden steel, to manufacture stainless steel, and to form many useful alloys. Much is used in plating to produce a hard, beautiful surface and to prevent corrosion. Chromium gives glass an emerald green color and is widely used as a catalyst.

The refractory industry has found chromite useful for forming bricks and shapes, as it has a high melting point, moderate thermal expansion, and stability of crystalline structure.

Compounds

All compounds of chromium are colored; the most important are the chromates of sodium and potassium and the dichromates and the potassium and ammonium chrome alums. The dichromates are used as oxidizing agents in quantitative analysis, also in tanning leather.

Other compounds are of industrial value; lead chromate is chrome yellow, a valued pigment. Chromium compounds are used in the textile industry as mordants, and by the aircraft and other industries for anodizing aluminum.

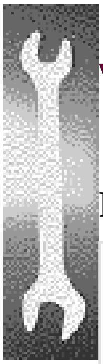
Handling

Chromium compounds are toxic and should be handled with proper safeguards.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Vanadium

For tools

Atomic Number:	23
Atomic Symbol:	V
Atomic Weight:	50.9414
Electron Configuration:	$[Ar]4s^2 3d^3$

History

(Scandinavian goddess, Vanadis) Vanadium was first discovered by del Rio in 1801. Unfortunately, a French chemist incorrectly declared that del Rio's new element was only impure chromium. Del Rio thought himself to be mistaken and accepted the French chemists's statement.

The element was rediscovered in 1830 by Sefstrom, who named the element in honor of the Scandinavian goddess, Vanadis, because of its beautiful multicolored compounds. It was isolated in nearly pure form by Roscoe, in 1867, who reduced the chloride with hydrogen.

Vanadium of 99.3 to 99.8% purity was not produced until 1922.

Sources

Vanadium is found in about 65 different minerals among which are carnotite, roscoelite, vanadinite, and patronite, important sources of the metal. Vanadium is also found in phosphate rock and certain iron ores, and is present in some crude oils in the form of organic complexes. It is also found in small percentages in meteorites.

Commercial production from petroleum ash holds promise as an important source of the element. High-purity ductile vanadium can be obtained by reduction of vanadium trichloride with magnesium or with magnesium-sodium mixtures.

Much of the vanadium metal being produced is now made by calcium reduction of V_2O_5 in a pressure vessel, an adaption of a process developed by McKechnie and Seybair.

Isotopes

Natural vanadium is a mixture of two isotopes, 50V (0.24%) and 51V (99.76%). 50V is slightly radioactive, having a half-life of > 3.9 x 10^{17} years. Nine other unstable isotopes are recognized.

Properties

Pure vanadium is a bright white metal, and is soft and ductile. It has good corrosion resistance to alkalis, sulfuric and hydrochloric acid, and salt water, but the metal oxidizes readily above 660^{o}C.

The metal has good structural strength and a low fission neutron cross section, making it useful in nuclear applications.

Uses

Vanadium is used in producing rust resistant and high speed tools steels. It is an important carbide stabilizer in making steels.

About 80% of the vanadium now produced is used as ferrovanadium or as a steel additive. Vanadium foil is used as a bonding agent in cladding titanium to steel. Vanadium pentoxide is used in ceramics and as a catalyst.

It is also used to produce a superconductive magnet with a field of 175,000 gauss.

Handling

Vanadium and its compounds are toxic and should be handled with care. The maximum allowable concentration of V_2O_5 dust in air is about 0.05 (8-hour time-weighted average - 40-hour week).

Costs

Ductile vanadium is commercially available. Commercial vanadium metal, of about 95% purity, costs about $20 / lb. Vanadium (99.9%) costs about $100 / oz.

Isotopes

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Yttrium

For color TV screens.

Atomic Number:	39
Atomic Symbol:	Y
Atomic Weight:	88.9059
Electron Configuration:	[Kr]$5s^1 4d^1$

History

(Ytterby, a village in Sweden near Vauxholm) Yttria, which is an earth containing yttrium, was discovered by Gadolin in 1794. Ytterby is the site of a quarry which yielded many unusual minerals containing rare earths and other elements. This small town, near Stockholm, bears the honor of giving names to erbium, terbium, and ytterbium as well as yttrium.

In 1843 Mosander showed that yttira could be resolved into the oxides (or earths) of three elements. The name yttria was reserved for the most basic one; the others were named erbia and terbia.

Sources

Yttrium occurs in nearly all of the rare-earth minerals. Analysis of lunar rock samples obtained during the Apollo missions show a relatively high yttrium content.

It is recovered commercially from monazite sand, which contains about 3%, and from bastnasite, which contains about 0.2%. Wohler obtained the impure element in 1828 by reduction of the anhydrous chloride with potassium. The metal is now produced commercially by reduction of the fluoride with calcium metal. It can also be prepared by other techniques.

Properties

Yttrium has a silver-metallic luster and is relatively stable in air. Turnings of the metal, however, ignite in air if their temperature exceeds 400_{o}C. Finely divided yttrium is very unstable in air.

Uses

Yttrium oxide is one of the most important compounds of yttrium and accounts for the largest use. It is widely used in making YVO_4 europium, and Y_2O_3 europium phosphors to give the red color in color television tubes. Many hundreds of thousands of pounds are now used in this application.

Yttrium oxide also is used to produce yttrium-iron-garnets, which are very effective microwave filters.

Yttrium iron, aluminum, and gadolinium garnets, with formulas such as $Y_3Fe_5O_{12}$ and $Y_3Al_5O_{12}$, have interesting magnetic properties. Yttrium iron garnet is also exceptionally efficient as both a transmitter and transducer of acoustic energy. Yttrium aluminum garnet, with a hardness of 8.5, is also finding use as a gemstone (simulated diamond).

Small amounts of yttrium (0.1 to 0.2%) can be used to reduce the grain size in chromium, molybdenum, zirconium, and titanium, and to increase strength of aluminum and magnesium alloys.

Alloys with other useful properties can be obtained by using yttrium as an additive. The metal can be used as a deoxidizer for vanadium and other nonferrous metals. The metal has a low cross section for nuclear capture. 90Y, one of the isotopes of yttrium, exists in equilibrium with its parent 90Sr, a product of nuclear explosions. Yttrium has been considered for use as a nodulizer for producing nodular cast iron, in which the graphite forms compact nodules instead of the usual flakes. Such iron has increased ductility.

Yttrium is also finding application in laser systems and as a catalyst for ethylene polymerization.

It also has potential use in ceramic and glass formulas, as the oxide has a high melting point and imparts shock resistance and low expansion characteristics to glass.

Isotopes

Natural yttrium contains but one isotope, 89Y. Nineteen other unstable isotopes have been characterized.

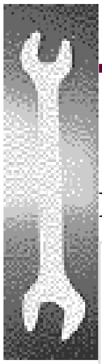
Costs

Yttrium metal of 99.9% purity is commercially available at a cost of about $75/oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Titanium

For expensive tools

Atomic Number:	22
Atomic Symbol:	Ti
Atomic Weight:	47.90
Electron Configuration:	$[Ar]4s^2 3d^2$

History

(L. titans, the first sons of the Earth, mythology)

Discovered by Gregor in 1791; named by Klaproth in 1795. Impure titanium was prepared by Nilson and Pettersson in 1887; however, the pure metal (99.9%) was not made until 1910 by Hunter by heating $TiCl_4$ with sodium in a steel bomb.

Sources

Titanium is present in meteorites and in the sun. Rocks obtained during the Apollo 17 lunar mission showed presence of 12.1% TiO2 and rocks obtained during earlier Apollo missions show lower percentages.

Titanium oxide bands are prominent in the spectra of M-type stars. The element is the ninth most abundant in the crust of the earth. Titanium is almost always present in igneous rocks and in the sediments derived from them.

It occurs in the minerals rutile, ilmenite, and sphene, and is present in titanates and in many iron ores. Titanium is present in the ash of coal, in plants, and in the human body.

The metal was a laboratory curiosity until Kroll, in 1946, showed that titanium could be produced commercially by reducing titanium tetrachloride with magnesium. This method is largely used for producing the metal today. The metal can be purified by decomposing the iodide.

Properties

Titanium, when pure, is a lustrous, white metal. It has a low density, good strength, is easily fabricated, and has excellent corrosion resistance. It is ductile only when it is free of oxygen. The metal, which burns in air, is the only element that burns in nitrogen.

Titanium is resistant to dilute sulfuric and hydrochloric acid, most organic acids, most chlorine gas, and chloride solutions.

Natural titanium is reported to become very radioactive after bombardment with deuterons. The emitted radiations are mostly positrons and hard gamma rays. The metal is dimorphic. The hexagonal alpha form changes to the cubic beta form very slowly at about 880$_o$C. The metal combines with oxygen at red heat, and with chlorine at 550$_o$C.

Titanium metal is considered to be physiologically inert. When pure, titanium dioxide is relatively clear and has an extremely high index of refraction with an optical dispersion higher than diamond.

Isotopes

Natural titanium consists of five isotopes with atomic masses from 46 to 50. All are stable. Eight other unstable isotopes are known.

Uses

Titanium is important as an alloying agent with aluminum, molybdenum, manganese, iron, and other metals. Alloys of titanium are principally used for aircraft and missiles where lightweight strength and ability to withstand extremes of temperature are important.

Titanium is as strong as steel, but 45% lighter. It is 60% heavier than aluminum, but twice as strong.

Titanium has potential use in desalination plants for converting sea water into fresh water. The metal has excellent resistance to sea water and is used for propeller shafts, rigging, and other parts of ships exposed to salt water. A titanium anode coated with platinum has been used to provide cathodic protection from corrosion by salt water.

It is produced artificially for use as a gemstone, but it is relatively soft. Star sapphires and rubies exhibit their asterism as a result of the presence of TiO2.

Titanium dioxide is extensively used for both house paint and artist's paint, because it is permanent and has good covering power. Titanium oxide pigment accounts for the largest use of the element. Titanium paint is an excellent reflector of infrared, and is extensively used in solar observatories where heat causes poor seeing conditions.

Titanium tetrachloride is used to iridize glass. This compound fumes strongly in air and has been used to produce smoke screens.

Costs

The price of titanium metal powder (99.95%) is about $100/lb.

Isotopes available at Los Alamos National Laboratory

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Molybdenum

For filament in electric heaters.

Atomic Number:	42
Atomic Symbol:	Mo
Atomic Weight:	95.94
Electron Configuration:	$[Kr]5s^1 4d^5$

History

(Gr. molybdos, lead) Before Scheele recognized molybdenite as a distinct ore of a new element in 1778, it was confused with graphite and lead ore. The metal was prepared as an impure form in 1782 by Hjelm. Molybdenum does not occur native, but is obtained principally from molybdenite. Wulfenite, and Powellite are also minor commercial ores.

Sources

Molybdenum is also recovered as a by-product of copper and tungsten mining operations. The metal is prepared from the powder made by the hydrogen reduction of purified molybdic trioxide or ammonium molybdate.

Properties

The metal is silvery white, very hard, but is softer and more ductile than tungsten. It has a high elastic modulus, and only tungsten and tantalum, of the more readily available metals, have higher melting points. It is a valuable alloying agent, as it contributes to the hardenability and toughness of quenched and tempered steels. It also improves the strength of steel at high temperatures.

Uses

It is used in certain nickel-based alloys, such as the "Hastelloys(R)" which are heat-resistant and corrosion-resistant to chemical solutions. Molybdenum oxidizes at elevated temperatures. The metal has found recent application as electrodes for electrically heated glass furnaces and foreheaths. The metal is also used in nuclear energy applications and for missile and aircraft parts. Molybdenum is valuable as a catalyst in the refining of petroleum. It has found applications as a filament material in electronic and electrical applications. Molybdenum is an

essential trace element in plant nutrition. Some lands are barren for lack of this element in the soil. Molybdenum sulfide is useful as a lubricant, especially at high temperatures where oils would decompose. Almost all ultra-high strength steels with minimum yield points up to 300,000 psi(lb / $in.^2$) contain molybdenum in amounts from 0.25 to 8%.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

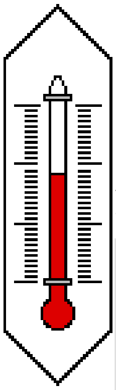
Last Updated: 12/19/97, CST Information Services Team

Common Isotopes

Isotope	Ci/g	Ci/L to ppm multiply by:	ppm to Ci/L multiply by:
Ac-227	72.9	1.37E+01	7.29E-02
Ac-228	2.24E+06	4.46E-04	2.24E+03
Am-241	3.24E+00	3.09E+02	3.24E-03
Am-242m	9.72E+00	1.03E+02	9.72E-03
Am-242	8.09E+05	1.24E-03	8.09E+02
Am-243	1.85E-01	5.41E+03	1.85E-04
Am-244	1.27E+06	7.87E-04	1.27E+03
Ba-131	8.31E+04	1.20E-02	8.31E+01
Ba-140	7.29E+04>	1.37E-02	7.29E+01
Cd-109	2.64E+03	3.79E-01	2.64E+00
Cd-115m	2.60E+04	3.85E-02	2.60E+01
Cd-115	5.09E+05	1.96E-03	5.09E+02
Cs-137	87	1.15E+01	8.70E-02
Cs-138	4.20E+07	2.38E-05	4.20E+04
Co-157	8480	1.18E-01	8.48E+00
Co-60	1133	8.83E-01	1.13E+00
Mo-99	4.75E+05	2.11E-03	4.75E+02
Np-237	7.05E-04	1.42E+06	7.05E-07
Np-239	2.33E+05	4.29E-03	2.33E+02
Pu-238	17.4	5.75E+01	1.74E-02
Pu-239	0.0614	1.63E+04	6.14E-05
Pu-240	0.226	4.42E+03	2.26E-04
Pu-241	112	8.93E+00	1.12E-01
Pu-242	3.90E-03	2.56E+05	3.90E-06
Pu-243	2.59E+06	3.86E-04	2.59E+03
Pu-244	1.90E-05	5.26E+07	1.90E-08
Sr-89	2.78E+04	3.60E-02	2.78E+01
Sr-90	143	6.99E+0	1.4E-01
U-235	2.10E-06	4.76E+08	2.10E-09
U-238	3.33E-07	3.00E+09	3.33E-10

Last Updated: 8/5/99

Periodic Table



Mercury

For thermometers.

Atomic Number:	80
Atomic Symbol:	Hg
Atomic Weight:	200.59
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^{10}$

History

(Planet Mercury) Known to ancient Chinese and Hindus; found in Egyptian tombs of 1500 B.C. Mercury is the only common metal liquid at ordinary temperatures. It only rarely occurs free in nature. The chief ore is cinnabar, Spain and Italy produce about 50% of the world's supply of the metal. The commercial unit for handling mercury is the "flask," which weighs 76 lb. The metal is obtained by heating cinnabar in a current of air and by condensing the vapor. It is a heavy, silvery-white metal; a rather poor conductor of heat, as compared with other metals, and a fair conductor of electricity. It easily forms alloys with many metals, such as gold, silver, and tin, which are called amalgams. Its ease in amalgamating with gold is made use of in the recovery of gold from its ores. The most important salts are mercury chloride (corrosive sublimate - a violent poison), mercurous chloride (calomel, occasionally still used in medicine), mercury fulminate, a detonator widely used in explosives, and mercuric sulfide (vermilion, a high-grade paint pigment). Organic mercury compounds are important. It has been found that an electrical discharge causes mercury vapor to combine with neon, argon, krypton, and xenon. These products, held together with van der Waals' forces, correspond to HgNe, HgAr, HgKr, and HgXe. Mercury is a virulent poison and is readily absorbed through the respiratory tract, the gastrointestinal tract, or through unbroken skin. It acts as a cumulative poison and dangerous levels are readily attained in air. Air saturated with mercury vapor at 20C contains a concentration that exceeds the toxic limit many times. The danger increases at higher temperatures. It is therefore important that mercury be handled with care. Containers of mercury should be securely covered and spillage should be avoided. It it is necessary to heat mercury or mercury compounds, it should be done in a well-ventilated hood. Methyl mercury is a dangerous pollutant and is now widely found in water and streams. The triple point of mercury, -38.8344C, is a fixed point on the International Temperature Scale (ITS-90).

Uses

The metal is widely used in laboratory work for making thermometers, barometers, diffusion pumps, and many other instruments. It is used in making mercury-vapor lamps and advertising

signs, etc. and is used in mercury switches and other electronic apparatus. Other uses are in making pesticides, cercury cells for caustic soda and chlorine production, dental preparations, anti-fouling paint, batteries, and catalysts.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Cobalt

For permanent magnets.

Atomic Number:	27
Atomic Symbol:	Co
Atomic Weight:	58.9332
Electron Configuration:	$[Ar]4s^2 3d^7$

History

(Kobald, from the German, goblin or evil spirit, cobalos, Greek, mine) Brandt discovered cobalt in about 1735.

Sources

Cobalt occurs in the minerals cobaltite, smaltite, and erythrite, and is often associated with nickel , silver , lead , copper , and iron ores, from which it is most frequently obtained as a by-product. It is also present in meteorites.

Important ore deposits are found in Zaire, Morocco, and Canada. The U.S. Geological Survey has announced that the bottom of the north central Pacific Ocean may have cobalt-rich deposits at relatively shallow depths in water close to the the Hawaiian Islands and other U.S. Pacific territories.

Properties

Cobalt is a brittle, hard metal, resembling iron and nickel in appearance. It has a metallic permeability of about two thirds that of iron. Cobalt tends to exist as a mixture of two allotropes over a wide temperature range. The transformation is sluggish and accounts in part for the wide variation in reported data on physical properties of cobalt.

Uses

It is alloyed with iron, nickel and other metals to make Alnico, an alloy of unusual magnetic strength with many important uses. Stellite alloys, containing cobalt, chromium , and tungsten , are used for high-speed, heavy-duty, high temperature cutting tools, and for dies.

Cobalt is also used in other magnetic steels and stainless steels, and in alloys used in jet turbines

and gas turbine generators. The metal is used in electroplating because of its appearance, hardness, and resistance to oxidation.

The salts have been used for centuries to produce brilliant and permanent blue colors in porcelain, glass, pottery, tiles, and enamels. It is the principal ingredient in Sevre's and Thenard's blue. A solution of the chloride is used as a sympathetic ink. Cobalt carefully used in the form of the chloride, sulfate, acetate, or nitrate has been found effective in correcting a certain mineral deficiency disease in animals.

Soils should contain 0.13 to 0.30 ppm of cobalt for proper animal nutrition.

Isotope

Cobalt-60, an artificial isotope, is an important gamma ray source, and is extensively used as a tracer and a radiotherapeutic agent.

Costs

Single compact sources of Cobalt-60 vary from about $1 to $10/curie, depending on quantity and specific activity.

Handling

Exposure to cobalt (metal fumes and dust) should be limited to 0.05 mg/m^3 (8-hour time-weighted average 40-hour week).

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Lead

For batteries.

Atomic Number:	82
Atomic Symbol:	Pb
Atomic Weight:	207.2
Electron Configuration:	$[Xe]6s^{2}4f^{14}5d^{10}6p^{2}$

History

(Anglo-Saxon lead; L. plumbum) Long known, mentioned in Exodus. The alchemists believed lead to be the oldest metal and associated with the planet Saturn. Native lead occurs in nature, but it is rare.

Sources

Lead is obtained chiefly from galena (PbS) by a roasting process. Anglesite, cerussite, and minim are other common lead minerals.

Properties

Lead is a bluish-white metal of bright luster, is very soft, highly malleable, ductile, and a poor conductor of electricity. It is very resistant to corrosion; lead pipes bearing the insignia of Roman emperors, used as drains from the baths, are still in service. It is used in containers for corrosive liquids (such as sulfuric acid) and may be toughened by the addition of a small percentage of antimony or other metals.

Forms

Natural lead is a mixture of four stable isotopes: 204Pb (1.48%), 206Pb (23.6%), 207Pb (22.6%), and 208Pb (52.3%). Lead isotopes are the end products of each of the three series of naturally occurring radioactive elements: 206Pb for the uranium series, 207Pb for the actinium series, and 208Pb for the thorium series. Twenty seven other isotopes of lead, all of which are radioactive, are recognized.

Its alloys include solder, type metal, and various antifriction metals. Great quantities of lead, both as the metal and as the dioxide, are used in storage batteries. Much metal also goes into cable covering, plumbing, ammunition, and in the manufacture of lead tetraethyl.

Uses

The metal is very effective as a sound absorber, is used as a radiation shield around X-ray equipment and nuclear reactors, and is used to absorb vibration. White lead, the basic carbonate, sublimed white lead, chrome yellow, and other lead compounds are used extensively in paints, although in recent years the use of lead in paints has been drastically curtailed to eliminate or reduce health hazards.

Lead oxide is used in producing fine "crystal glass" and "flint glass" of a high index of refraction for achromatic lenses. The nitrate and the acetate are soluble salts. Lead salts such as lead arsenate have been used as insecticides, but their use in recent years has been practically eliminated in favor of less harmful organic compounds.

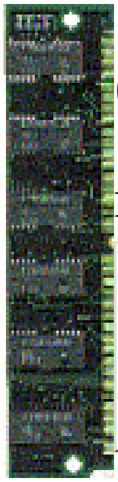
Handling

Care must be used in handling lead as it is a cumulative poison. Environmental concerns with lead poisoning has resulted in a national program to eliminate the lead in gasoline.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Gallium

For computer memory.

Atomic Number:	31
Atomic Symbol:	Ga
Atomic Weight:	69.72
Electron Configuration:	$[Ar]4s^{2}3d^{10}4p^{1}$

History

(L. Gallia, France; also from Latin, gallus, a translation of Lecoq, a cock) Predicted and described by Mendeleev as ekaaluminum, and discovered spectroscopically by Lecoq de Boisbaudran in 1875, who in the same year obtained the free metal by electrolysis of a solution of the hydroxide in KOH.

Sources

Gallium is often found as a trace element in diaspore, sphalerite, germanite, bauxite, and coal. Some flue dusts from burning coal have been shown to contain as much 1.5 percent gallium.

Properties

It is one of four metals -- mercury , cesium , and rubidium -- which can be liquid near room temperature and, thus, can be used in high-temperature thermometers. It has one of the longest liquid ranges of any metal and has a low vapor pressure even at high temperatures.

There is a strong tendency for gallium to supercool below its freezing point. Therefore, seeding may be necessary to initiate solidification.

Ultra-pure gallium has a beautiful, silvery appearance, and the solid metal exhibits a conchoidal fracture similar to glass. The metal expands 3.1 percent on solidifying; therefore, it should not be stored in glass or metal containers, because they may break as the metal solidifies.

High-purity gallium is attacked only slowly by mineral acids.

Uses

Gallium wets glass or porcelain and forms a brilliant mirror when it is painted on glass. It is widely used in doping semiconductors and producing solid-state devices such as transistors.

Magnesium gallate containing divalent impurities, such as Mn+2, is finding use in commercial ultraviolet-activated powder phosphors. Gallium arsenide is capable of converting electricity directly into coherent light. Gallium readily alloys with most metals, and has been used as a component in low-melting alloys.

Handling

Its toxicity appears to be of a low order, but should be handled with care until more data is available.

Costs

The metal can be supplied in ultra pure form (99.99999+%). The cost is about $3/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Cesium

For photoelectric cells.

Atomic Number:	55
Atomic Symbol:	Cs
Atomic Weight:	132.9054
Electron Configuration:	$[Xe]6s^1$

History

(L. caesius, sky blue) Cesium was discovered spectroscopically by Bunsen and Kirchhoff in 1860 in mineral water from Durkheim.

Sources

Cesium, an alkali metal, occurs in lepidolite, pollucte (a hydrated silicate of aluminum and cesium), and in other sources. One of the world's richest sources of cesium is located at Bernic Lake, Manitoba. The deposits are estimated to contain 300,000 tons of pollucite, averaging 20% cesium.

It can be isolated by elecytrolysis of the fused cyanide and by a number of other methods. Very pure, gas-free cesium can be prepared by thermal decomposition of cesium azide.

Properties

The metal is characterized by a spectrum containing two bright lines in the blue along with several others in the red, yellow, and green. It is silvery white, soft, and ductile. It is the most electropositive and most alkaline element.

Cesium, gallium, and mercury are the only three metals that are liquid at room temperature. Cesium reacts explosively with cold water, and reacts with ice at temperatures above -116C. Cesium hydroxide, the strongest base known, attacks glass.

Uses

Because of it has great affinity for oxygen, the metal is used as a "getter" in electron tubes. It is also used in photoelectric cells, as well as a catalyst in the hydrogenation of certain organic compounds.

The metal has recently found application in ion propulsion systems. Cesium is used in atomic clocks, which are accurate to 5 s in 300 years. Its chief compounds are the chloride and the nitrate.

Isotope

Cesium has more isotopes than any element--32--with masses ranging from 114 to 145.

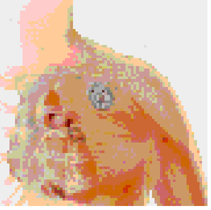
Costs

The present price of cesium is about $30/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Rubidium

For heart muscle research.

Atomic Number:	37
Atomic Symbol:	Rb
Atomic Weight:	85.4678
Electron Configuration:	$[Kr]5s^1$

History

(L. rubidus, deepest red) Discovered in 1861 by Bunsen and Kirchoff in the mineral lepidolite by use of the spectroscope.

Sources

The element is much more abundant than was thought several years ago. It is now considered to be the 16th most abundant element in the earth's crust. Rubidium occurs in pollucite, leucite, and zinnwaldite, which contains traces up to 1%, in the form of the oxide. It is found in lepidolite to the extent of about 1.5%, and is recovered commercially from this source. Potassium minerals, such as those found at Searles Lake, California, and potassium chloride recovered from the brines in Michigan also contain the element and are commercial sources. It is also found along with cesium in the extensive deposits of pollucite at Bernic Lake, Manitoba.

Properties

Rubidium can be liquid at room temperature. It is a soft, silvery-white metallic element of the alkali group and is the second most electropositive and alkaline element. It ignites spontaneously in air and reacts violently in water, setting fire to the liberated hydrogen. As with other alkali metals, it forms amalgams with mercury and it alloys with gold, cesium, sodium, and potassium. It colors a flame yellowish violet. Rubidium metal can be prepared by reducing rubidium chloride with calcium, and by a number of other methods. It must be kept under a dry mineral oil or in a vacuum or inert atmosphere.

Isotopes

Twenty four isotopes of rubidium are known. Naturally occurring rubidium is made of two isotopes, 85Rb and 87Rb. Rubidium-87 is present to the extent of 27.85% in natural rubidium and

is a beta emitter with a half-life of 4.9 x 10^{10} years. Ordinary rubidium is sufficiently radioactive to expose a photographic film in about 30 to 60 days. Rubidium forms four oxides: Rb_2O, Rb_2O_2, Rb_2O_3, Rb_2O_4.

Uses

Because rubidium can be easily ionized, it has been considered for use in "ion engines" for space vehicles; however, cesium is somewhat more efficient for this purpose. It is also proposed for use as a working fluid for vapor turbines and for use in a thermoelectric generator using the magnetohydrodynamic principle where rubidium ions are formed by heat at high temperature and passed through a magnetic field. These conduct electricity and act like an amature of a generator thereby generating an electric current. Rubidium is used as a getter in vacuum tubes and as a photocell component. It has been used in making special glasses. $RbAg_4I_5$ is important, as it has the highest room conductivity of any known ionic crystal. At 20°C its conductivity is about the same as dilute sulfuric acid. This suggests use in thin film batteries and other applications.

Cost

The present cost in small quantities is about $25/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Germanium

For wide-angle lenses.

Atomic Number:	32
Atomic Symbol:	Ge
Atomic Weight:	72.59
Electron Configuration:	[Ar]$4s^2 3d^{10} 4p^2$

History

(Latin Germania, Germany) Mendeleev predicted the existence of Germanium in 1871 as ekasilicon, and the element was discovered by Winkler in 1886.

Sources

The metal is found in

- argyrodite, a sulfide of germanium and silver ;
- germanite, which contains 8 percent of the element;
- zinc ores;
- coal; and
- other minerals.

The element is commercially obtained from the dusts of smelters processing zinc ores, as well as recovered from combustion by-products of certain coals. A large reserve of the elements for future uses in insured in coal sources.

Germanium can be separated from other metals by fractional distillation of its volatile tetrachloride. The techniques permit the production of germanium of ultra-high purity.

Properties

The element is a gray-white metalloid. In its pure state, the element is crystalline and brittle, retaining its luster in air at room temperature. It is a very important semiconductor material. Zone-refining techniques have led to production of crystalline germanium for semiconductor use with an impurity of only one part in 10^{10}.

Uses

When germanium is doped with arsenic , gallium , or other elements, it is used as a transistor element in thousands of electronic applications. The most common use of germanium is as a semiconductor. Germanium is also finding many other applications including use as an alloying agent, as a phosphor in fluorescent lamps, and as a catalyst.

Germanium and germanium oxide are transparent to the infrared and are used in infrared spectroscopes and other optical equipment, including extremely sensitive infrared detectors.

The high index of refraction and dispersion properties of its oxide's have made germanium useful as a component of wide-angle camera lenses and microscope objectives.

The field of organogermanium chemistry is becoming increasingly important. Certain germanium compounds have a low mammalian toxicity, but a marked activity against certain bacteria, which makes them useful as chemotherapeutic agents.

Costs

The cost of germanium is about $3/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Arsenic

For LEDs.

Atomic Number:	33
Atomic Symbol:	As
Atomic Weight:	74.9216
Electron Configuration:	$[Ar]4s^2 3d^{10} 4p^3$

History

(L. arsenicum, Gr. arsenikon, yellow orpiment, identified with arenikos, male, from the belief that metals were different sexes; Arabic, Az-zernikh, the orpiment from Persian zerni-zar, gold) Elemental arsenic occurs in two solid modifications: yellow, and gray or metallic, with specific gravities of 1.97, and 5.73, respectively. It is believed that Albertus Magnus obtained the element in 1250 A.D. In 1649 Schroeder published two methods of preparing the element. Mispickel, arsenopyrite, (FeSAs) is the most common mineral from which, on heating, the arsenic sublimes leaving ferrous sulfide.

Properties

The element is a steel gray, very brittle, crystalline, semimetallic solid; it tarnishes in air, and when heated is rapidly oxidized to arsenous oxide with the odor of garlic. Arsenic and its compounds are poisonous.

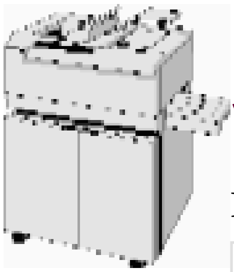
Uses

Arsenic is used in bronzing, pyrotechny, and for hardening and improving the sphericity of shot. The most important compounds are white arsenic, the sulfide, Paris green, calcium arsenate, and lead arsenate; the last three have been used as agricultural insecticides and poisons. Marsh's test makes use of the formation and ready decomposition of arsine. Arsenic is finding increasing uses as a doping agent in solid-state devices such as transistors. Gallium arsenide is used as a laser material to convert electricity directly into coherent light.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Selenium

For copy machines.

Atomic Number:	34
Atomic Symbol:	Se
Atomic Weight:	78.96
Electron Configuration:	$[Ar]4s^2 3d^{10} 4p^4$

History

(Gr. Selene, moon) Discovered by Berzelius in 1817, who found it associated with tellurium, named for the earth.

Production

Selenium is found in a few rare minerals such as crooksite and clausthalite. In years past it has been obtained from flue dusts remaining from processing copper sulfide ores, but the anode metal from electrolytic copper refineries now provide the source of most of the world's selenium. Selenium is recovered by roasting the muds with soda or sulfuric acid, or by smelting them with soda and niter.

Properties

Selenium exists in several allotropic forms. Three are generally recognized, but as many as that have been claimed. Selenium can be prepared with either an amorphous or crystalline structure. The color of amorphous selenium is either red, in powder form, or black, in vitreous form. Crystalline monoclinic selenium is a deep red; crystalline hexagonal selenium, the most stable variety, is a metallic gray.

Selenium exhibits both photovoltaic action, where light is converted directly into electricity, and photoconductive action, where the electrical resistance decreases with increased illumination. These properties make selenium useful in the production of photocells and exposure meters for photographic use, as well as solar cells. Selenium is also able to convert a.c. electricity to d.c., and is extensively used in rectifiers. Below its melting point selenium is a p-type semiconductor and is finding many uses in electronic and solid-state applications.

Elemental selenium has been said to be practically nontoxic and is considered to be an essential trace element; however, hydrogen selenide and other selenium compounds are extremely toxic, and resemble arsenic in their physiological reactions.

Isotopes

Naturally selenium contains six stable isotopes. Fifteen other isotopes have been characterized. The element is a member of the sulfur family and resembles sulfur both in its various forms and in its compounds.

Uses

Selenium is used in Xerography for reproducing and copying documents, letters, etc. It is used by the glass industry to decolorize glass and to make ruby-colored glasses and enamels. It is also used as a photographic toner, and as an additive to stainless steel.

Handling

Hydrogen selenide in a concentration of 1.5 ppm is intolerable to man. Selenium occurs in some solid in amounts sufficient to produce serious effects on animals feeding on plants, such as locoweed, grown in such soils. Exposure to selenium compounds (as Se) in air should not exceed 0.2 mg/m^3 (8-hour time-weighted average - 40-hour week).

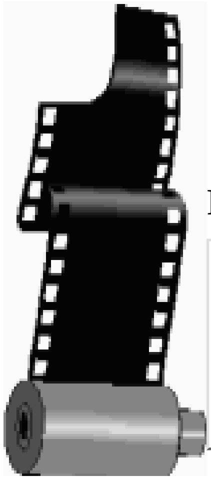
Cost

Selenium is priced at about $300/lb. It is also available in high-purity form at a somewhat higher cost.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Bromine

For photographic film.

Atomic Number:	35
Atomic Symbol:	Br
Atomic Weight:	79.904
Electron Configuration:	$[Ar]4s^2 3d^{10} 4p^5$

History

(Gr. bromos, stench) Discovered by Balard in 1826, but not prepared in quantity until 1860.

Sources

A member of the halogen group of elements, it is obtained from natural brines from wells in Michigan and Arkansas. Little bromine is extracted today from seawater, which contains only about 85 ppm.

Properties

Bromine is the only liquid nonmetallic element. It is a heavy, mobile, reddish-brown liquid, volatilizing readily at room temperature to a red vapor with a strong disagreeable odor, resembling chlorine, and having a very irritating effect on the eyes and throat; it is readily soluble in water or carbon disulfide, forming a red solution, is less active than chlorine but more so than iodine; it unites readily with many elements and has a bleaching action; when spilled on the skin it produces painful sores. It presents a serious health hazard, and maximum safety precautions should be taken when handling it.

Production

Much of the bromine output in the U.S. was used in the production of ethylene dibromide, a lead scavenger used in making gasoline antiknock compounds. Lead in gasoline, however, has been drastically reduced, due to environmental considerations. This will greatly affect future production of bromine.

Uses

Bromine is used in making fumigants, flameproofing agents, water purification compounds, dyes, medicinals, sanitizers, inorganic bromides for photography, etc. Organic bromides are also important.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Krypton

For fluorescent bulbs.

Atomic Number:	36
Atomic Symbol:	Kr
Atomic Weight:	83.80
Electron Configuration:	$[Ar]4s^2 3d^{10} 4p^6$

History

(Gr. kryptos, hidden) Discovered in 1898 by Ramsay and Travers in the residue left after liquid air had nearly boiled away. In 1960 it was internationally agreed that the fundamental unit of length, the meter, should be defined in terms of the orange-red spectral line of 86Kr. This replaced the standard meter of Paris, which was defined in terms of a bar made of a platinum-iridium alloy. In October 1983 the meter, which originally was defined as being one ten millionth of a quadrant of the earth's polar circumference, was again redefined by the International Bureau of Weights and Measures as being the length of a path traveled by light in a vacuum during a time interval of 1/299,792,458 of a second.

Sources

Krypton is present in the air to the extent of about 1 ppm. The atmosphere of Mars has been found to contain 0.3 ppm of krypton. Solid krypton is a white crystalline substance with a face-centered cubic structure which is common to all the "rare gases."

Properties

It is one of the "noble" gases. It is characterized by its brilliant green and orange spectral lines.

Isotopes

Naturally occurring krypton contains six stable isotopes. Seventeen other unstable isotopes are now recognized. The spectral lines of krypton are easily produced and some are very sharp. While krypton is generally thought of as a rare gas that normally does not combine with other elements to form compounds, it now appears that the existence of some krypton compounds is established. Krypton difluoride has been prepared in gram quantities and can be made by several methods. A higher fluoride of krypton and a salt of an oxyacid of krypton also have been

reported. Molecule-ions of ArKr+ and KrH+ have been identified and investigated, and evidence is provided for the formation of KrXe or KrXe+.

Uses

Krypton clathrates have been prepared with hydroquinone and phenol. 85Kr has found recent application in chemical analysis. By imbedding the isotope in various solids, kryptonates are formed. The activity of these kryptonates is sensitive to chemical reactions at the surface. Estimates of the concentration of reactants are therefore made possible. Krypton is used in certain photographic flash lamps for high-speed photography. Uses thus far have been limited because of its high cost. Krypton gas presently costs about $30/l.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Strontium

For nuclear batteries in buoys.

Atomic Number:	38
Atomic Symbol:	Sr
Atomic Weight:	87.62
Electron Configuration:	$[Kr]5s^2$

History

(Strontian, town in Scotland) Isolated by Davey by electrolysis in 1808; however, Adair Crawford in 1790 recognized a new mineral (strontianite) as differing from other barium minerals.

Forms

Strontium is found chiefly as celestite and strontianite. The metal can be prepared by electrolysis of the fused chloride mixed with potassium chloride, or is made by reducing strontium oxide with aluminum in a vacuum at a temperature at which strontium distills off. Three allotropic forms of the metal exist, with transition points at 235 and 540_oC.

Properties

Strontium is softer than calcium and decomposes in water more vigorously. It does not absorb nitrogen below 380_oC. It should be kept under kerosene to prevent oxidation. Freshly cut strontium has a silvery appearance, but rapidly turns a yellowish color with the formation of the oxide. The finely divided metal ignites spontaneously in air. Volatile strontium salts impart a beautiful crimson color to flames, and these salts are used in pyrotechnics and in the production of flares. Natural strontium is a mixture of four stable isotopes.

Isotopes

Sixteen other unstable isotopes are known to exist. Of greatest importance is 90Sr with a half-life of 29 years. It is a product of nuclear fallout and presents a health problem. This isotope is one of the best long-lived high-energy beta emitters known, and is used in SNAP (Systems for Nuclear Auxilliary Power) devices. These devices hold promise for use in space vehicles, remote weather stations, navigational buoys, etc., where a lightweight, long-lived, nuclear-electric power source

is needed.

Uses

The major use for strontium at present is in producing glass for color television picture tubes. It has also found use in producing ferrite magnets and in refining zinc. Strontium titanate is an interesting optical material as it has an extremely high refractive index and an optical dispersion greater than that of diamond. It has been used as a gemstone, but is very soft. It does not occur naturally.

Cost

Strontium metal (98% pure) in January 1990 cost about $5/oz.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Niobium

For welding rods.

Atomic Number:	41
Atomic Symbol:	Nb
Atomic Weight:	92.9064
Electron Configuration:	[Kr]$5s^1 4d^4$

History

(Niobe, daughter of Tantalus) Discovered in 1801 by Hatchett in an ore sent to England more than a century before by John Winthrop the Younger, first governor of Connecticut. The metal was first prepared in 1864 by Blomstrand, who reduced the chloride by heating it in a hydrogen atmosphere. The name niobium was adopted by the International Union of Pur and Applied Chemicstry in 1950 after 100 years of controversy. Many leading chemical societies and government organizations refer to it by this name. Most metallurgists, leading metal societies, and all but one of the leading U.S. commercial producers, however, still refer to the metal as "columbium."

Sources

The element is found in niobite (or columbite), niobite-tantalite, parochlore, and euxenite. Large deposits of niobium have been found associated with carbonatites (carbon-silicate rocks), as a constituent of parochlore. Extensive ore reserves are found in Canada, Brazil, Nigeria, Zaire, and in Russia.

Properties

It is a shiny, white, soft, and ductile metal, and takes on a bluish cast when exposed to air at room temperatures for a long time. The metal starts to oxidize in air at 200_oC, and when processed at even moderate temperatures must be placed in a protective atmosphere.

Uses

It is used in arc-welding rods for stabilized grades of stainless steel. Thousands of pounds of niobium have been used in advanced air frame systems such as were used in the Gemini space program. The element has superconductive properties; superconductive magnets have been

made with Nb-Zr wire, which retains its superconductivity in strong magnetic fields. This type of application offers hope of direct large-scale generation of electric power.

Isotopes

Eighteen isotopes of niobium are known. The metal can be isolated from tantalum, and prepared in several ways.

Cost

Niobium metal (99.5% pure) is priced at about $75/lb.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Technetium

For radiation source for medical research.

Atomic Number:	43
Atomic Symbol:	Tc
Atomic Weight:	97
Electron Configuration:	[Kr]$5s^2 4d^5$

History

(Gr. technetos, artificial) Element 43 was predicted on the basis of the periodic table, and was erroneously reported as having been discovered in 1925, at which time it was named masurium. The element was actually discovered by Perrier and Segre in Italy in 1937. It was found in a sample of molybdenum, which was bombarded by deuterons in the Berkeley cyclotron, and which E. Lawrence sent to these investigators. Technetium was the first element to be produced artificially. Since its discovery, searches for the element in terrestrial material have been made. Finally in 1962, technetium-99 was isolated and identified in African pitchblende (a uranium rich ore) in extremely minute quantities as a spontaneous fission product of uranium-238 by B.T. Kenna and P.K. Kuroda. If it does exist, the concentration must be very small. Technetium has been found in the spectrum of S-, M-, and N-type stars, and its presence in stellar matter is leading to new theories of the production of heavy elements in the stars.

Isotopes

Twenty-two isotopes of technetium with masses ranging from 90 to 111 are reported. All the isotopes of technetium are radioactive. It is one of two elements with $Z < 83$ that have no stable isotopes; the other element is promethium ($Z = 61$). Technetium has three long lived radioactive isotopes: ^{97}Tc ($T_{1/2}$ =2.6 x 10^6 years), ^{98}Tc ($T_{1/2}$ = 4.2 x 10^6 years) and ^{99}Tc ($T_{1/2}$ = 2.1 x 10^5 years). ^{95}Tc$_m$ ("m" stands for meta state) ($T_{1/2}$ = 61 days) is used in tracer work. However, the most useful isotope of technetium is ^{99}Tc$_m$ ($T_{1/2}$ = 6.01 hours) is used in many medical radioactive isotope tests because of its half-life being short, the energy of the gamma ray it emits, and the ability of technetium to be chemically bound to many biologically active molecules. Because ^{99}Tc is produced as a fission product from the fission of uranium in nuclear reactors, large quantities have been produced over the years. There are kilogram quantities of technetium.

Properties

Technetium is a silvery-gray metal that tarnishes slowly in moist air. The common oxidation states of technetium are +7, +5, and +4. Under oxidizing conditions technetium (VII) will exist as the pertechnetate ion, TcO_4^-. The chemistry of technetium is said to be similar to that of rhenium. Technetium dissolves in nitric acid, aqua regia, and conc. sulfuric acid, but is not soluble in hydrochloric acid of any strength. The element is a remarkable corrosion inhibitor for steel. The metal is an excellent superconductor at 11K and below.

Cost

Until 1960, technetium was available only in small amounts and the price was as high as $2800/g. It is now commercially available to holders of O.R.N.L. permits at a price of $60/g.

Handling

It is reported that mild carbon steels may be effectively protected by as little as 55 ppm of $KTcO_4$ in aerated distilled water at temperatures up to 250°C. This corrosion protection is limited to closed systems, since technetium is radioative and must be confined. ^{98}Tc has a specific activity of 6.2×10^8 Bq/g. Activity of this level must not be allowed to spread. ^{99}Tc is a contamination hazard and should be handled in a glove box.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Ruthenium

For eye treatment.

Atomic Number:	44
Atomic Symbol:	Ru
Atomic Weight:	101.07
Electron Configuration:	$[Kr]5s^1 4d^7$

History

(L. Ruthenia, Russia) Berzelius and Osann in 1827 examined the residues left after dissolving crude platinum from the Ural mountains in aqua regia. While Berzelius found no unusual metals, Osann thought he found three new metals, one of which he named ruthenium. In 1844 Klaus, generally recognized as the discoverer, showed that Osann's ruthenium oxide was very impure and that it contained a new metal. Klaus obtained 6 g of ruthenium from the portion of crude platinum that is insoluble in aqua regia.

Sources

A member of the platinum group, ruthenium occurs native with other members of the group in ores found in the Ural mountains and in North and South America. It is also found along with other platinum metals in small but commercial quantities in pentlandite of the Sudbury, Ontario, nickel-mining region, and in pyroxinite deposits of South Africa.

Production

The metal is isolated commercially by a complex chemical process, the final stage of which is the hydrogen reduction of ammonium ruthenium chloride, which yields a powder. The powder is consolidated by powder metallurgy techniques or by argon-arc welding.

Properties

Ruthenium is a hard, white metal and has four crystal modifications. It does not tarnish at room temperatures, but oxidizes explosively. It is attacked by halogens, hydroxides, etc. Ruthenium can be plated by electrodeposition or by thermal decomposition methods. The metal is one of the most effective hardeners for platinum and palladium, and is alloyed with these metals to make electrical contacts for severe wear resistance. A ruthenium-molybdenum alloy is said to be

superconductive at 10.6 K. The corrosion resistance of titanium is improved a hundredfold by addition of 0.1% ruthenium. It is a versatile catalyst. Hydrogen sulfide can be split catalytically by light using an aqueous suspension of CdS particles loaded with ruthenium dioxide. It is thought this may have application to removal of H2S from oil refining and other industrial processes. Compounds in at least eight oxidation states have been found, but of these, the +2, +3, and +4 states are the most common. Ruthenium tetroxide, like osmium tetroxide, is highly toxic. In addition, it may explode. Ruthenium compounds show a marked resemblance to those of cadmium.

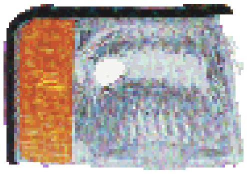
Cost

The metal is priced at about $30/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Rhodium

For headlight reflectors.

Atomic Number:	45
Atomic Symbol:	Rh
Atomic Weight:	102.9055
Electron Configuration:	$[Kr]5s^1 4d^8$

History

(Gr. rhodon, rose) Wollaston discovered rhodium in 1803-4 in crude platinum ore he presumably obtained from South America.

Sources

Rhodium occurs native with other platinum metals in river sands of the Urals and in North and South America. It is also found with other platinum metals in the copper-nickel sulfide area of the Sudbury, Ontario region. Although the quantity occurring here is very small, the large tonnages of nickel processed make the recovery commercially feasible. The annual world production of rhodium is only 7 or 8 tons.

Properties

The metal is silvery white and at red heat slowly changes in air to the resquioxide. At higher temperatures it converts back to the element. Rhodium has a higher melting point and lower density than platinum. It has a high reflectance and is hard and durable.

Uses

Rhodiums primary use is as an alloying agent to harden platinum and palladium. Such alloys are used for furnace windings, thermocouple elements, bushings for glass fiber production, electrodes for aircraft spark plugs, and laboratory crucibles. It is useful as an electrical contact material as it has a low electrical resistance, a low and stable contact resistance, and is highly resistant to corrosion. Plated rhodium, produced by electroplating or evaporation, is exceptionally hard and is used for optical instruments. Rhodium is also used for jewelry, for decoration, and as a catalyst.

Handling

Exposure to rhodium (metal fume and dust, as Rh) should not exceed 1 mg/m^3 (8-hour time-weighted average, 40-hour week).

Cost

Rhodium costs about $1,000/troy oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Palladium

For dental crowns.

Atomic Number:	46
Atomic Symbol:	Pd
Atomic Weight:	106.4
Electron Configuration:	$[Kr]4d^{10}$

History

Palladium was named after the asteroid Pallas, which was discovered at about the same time. Pallas was the Greek goddess of wisdom.

Sources

Discovered in 1803 by Wollaston, Palladium is found with platinum and other metals of the platinum group in placer deposits of Russia, South America, North America, Ethiopia, and Australia. It is also found associated with the nickel-copper deposits of South Africa and Ontario. Palladium's separation from the platinum metals depends upon the type of ore in which it is found.

Properties

The element is a steel-white metal, it does not tarnish in air, and it is the least dense and lowest melting of the platinum group of metals. When annealed, it is soft and ductile; cold-working greatly increases its strength and hardness. Palladium is attacked by nitric and sulfuric acid.

At room temperatures, the metal has the unusual property of absorbing up to 900 times its own volume of hydrogen, possibly forming Pd2H. It is not yet clear if this is a true compound. Hydrogen readily diffuses through heated palladium, providing a means of purifying the gas.

Uses

Finely divided palladium is a good catalyst and is used for hydrogenation and dehydrogenation reactions. It is alloyed and used in jewelry trades.

White gold is an alloy of gold decolorized by the addition of palladium. Like gold, palladium can be beaten into leaf as thin as 1/250,000 in. The metal is used in dentistry, watchmaking, and

in making surgical instruments and electrical contacts.

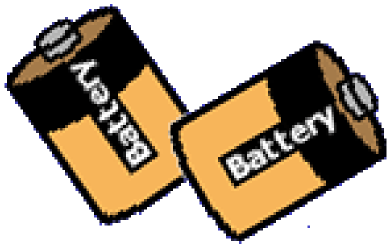
Cost

The metal sells for about $150 / troy oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Cadmium

For rechargeable batteries.

Atomic Number:	48
Atomic Symbol:	Cd
Atomic Weight:	112.40
Electron Configuration:	$[Kr]5s^2 4d^{10}$

History

(L. cadmia; Gr. kadmeia - ancient name for calamine, zinc carbonate) Discovered by Stromeyer in 1817 from an impurity in zinc carbonate. Cadmium most often occurs in small quantities associated with zinc ores, such as sphalerite (ZnS). Greenockite (CdS) is the only mineral of any consequence bearing cadmium. Almost all cadmium is obtained as a by-product in the treatment of zinc, copper, and lead ores. It is a soft, bluish-white metal which is easily cut with a knife. It is similar in many respects to zinc. Failure to appreciate the toxic properties of cadmium may cause workers to be unwittingly exposed to dangerous fumes. Silver solder, for example, which contains cadmium, should be handled with care. Serious toxicity problems have been found from long-term exposure and work with cadmium plating baths. Exposure to cadmium dust should not exceed 0.01 mg/m^3 (8-hour time-weighted average, 40-hour week). The ceiling concentration (maximum), for a period of 15 min, should not exceed 0.14 mg/m^3. Cadmium oxide fume exposure (8-hour, 40-hour week) should not exceed 0.05 mg/m^3, and the maximum concentration should not exceed 0.05 mg/m^3. These values are presently being restudied and recommendations have been made to reduce the exposure. In 1927 the International Conference on Weights and Measures redefined the meter in terms of the wavelength of the red cadmium spectral line (i.e. 1m = 1.553,164.13 wavelengths). This definition has been changed (see under Krypton).

Uses

Cadmium is a component of some of the lowest melting alloys; it is used in bearing alloys with low coefficients of friction and great resistance to fatigue; it is used extensively in electroplating, which accounts for about 60% of its use. It is also used in many types of solder, for standard E.M.F. cells, for Ni-Cd batteries, and as a barrier to control nuclear fission. Cadmium compounds are used in black and white television phosphors and in blue and green phosphors for color TV tubes. It forms a number of salts, of which the sulfate is most common; the sulfide is used as a yellow pigment. Cadmium and solutions of its compounds are toxic.

Cost

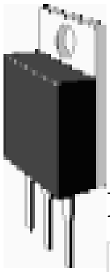
The current price of cadmium is about $12/lb. It is available in high purity form.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Indium

For transistors.

Atomic Number:	49
Atomic Symbol:	In
Atomic Weight:	114.82
Electron Configuration:	$[Kr]5s^2 4d^{10} 5p^1$

History

(from the brilliant indigo line in its spectrum) Discovered by Reich and Richter, who later isolated the metal. Until 1924, a gram or so constituted the world's supply of this element in isolated form. It is probably about as abundant as silver. About 4 million troy ounces of indium are now produced annually in the Free World. Canada is presently producing more than 1,000,000 troy ounces annually.

Sources

Indium is most frequently associated with zinc materials, and it is from these that most commercial indium is now obtained; however, it is also found in iron, lead, and copper ores.

Cost

The present cost of indium is about $1 to $5/g, depending on quantity and purity.

Properties

It is available in ultra pure form. Indium is a very soft, silvery-white metal with a brilliant luster. The pure metal gives a high-pitched "cry" when bent. It wets glass, as does gallium.

Uses

It has found application in making low-melting allows; an allow of 24% indium - 76% gallium is liquid at room temperature. It is used in making bearing alloys, germanium transistors, rectifiers, thermistors, and photoconductors. It can be plated onto metal and evaporated onto glass, forming a mirror as good as that made with silver but with more resistance to atmospheric corrosion.

Handling

There is evidence that indium has a low order of toxicity; however, care should be taken until further information is available.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Tin

For organ pipes.

Atomic Number:	50
Atomic Symbol:	Sn
Atomic Weight:	118.69
Electron Configuration:	$[Kr]5s^{2}4d^{10}5p^{2}$

History

(anglo-Saxon, tin; L. stannum) Known to the ancients.

Sources

Tin is found chiefly in cassiterite (SnO_2). Most of the world's supply comes from Malaya, Bolivia, Indonesia, Zaire, Thailand, and Nigeria. The U.S. produces almost none, although occurrences have been found in Alaska and California. Tin is obtained by reducing the ore with coal in a reverberatory furnace.

Properties

Ordinary tin is composed of nine stable isotopes; 18 unstable isotopes are also known. Ordinary tin is a silver-white metal, is malleable, somewhat ductile, and has a highly crystalline structure. Due to the breaking of these crystals, a "tin cry" is heard when a bar is bent.

Forms

The element has two allotropic forms at normal pressure. On warming, gray, or alpha tin, with a cubic structure, changes at $13.2^{o}C$ into white, or beta tin, the ordinary form of the metal. White tin has a tetragonal structure. When tin is cooled below $13.2^{o}C$, it changes slowly from white to gray. This change is affected by impurities such as aluminum and zinc, and can be prevented by small additions of antimony or bismuth. This change from the alpha to beta form is called the tin pest. There are few if any uses for gray tin. Tin takes a high polish and is used to coat other metals to prevent corrosion or other chemical action. Such tin plate over steel is used in the so-called tin can for preserving food.

Alloys of tin are very important. Soft solder, type metal, fusible metal, pewter, bronze, bell metal, Babbitt metal, White metal, die casting alloy, and phosphor bronze are some of the

important alloys using tin.

Tin resists distilled sea and soft tap water, but is attacked by strong acids, alkalis, and acid salts. Oxygen in solution accelerates the attack. When heated in air, tin forms Sn_2, which is feebly acid, forming stannate salts with basic oxides. The most important salt is the chloride, which is used as a reducing agent and as a mordant in calico printing. Tin salts sprayed onto glass are used to produce electrically conductive coatings. These have been used for panel lighting and for frost-free windshields. Most window glass is now made by floating molten glass on molten tin (float glass) to produce a flat surface (Pilkington process).

Of recent interest is a crystalline tin-niobium alloy that is superconductive at very low temperatures. This promises to be important in the construction of superconductive magnets that generate enormous field strengths but use practically no power. Such magnets, made of tin-niobium wire, weigh but a few pounds and produce magnetic fields that, when started with a small battery, are comparable to that of a 100 ton electromagnet operated continuously with a large power supply.

Handling

The small amount of tin found in canned foods is quite harmless. The agreed limit of tin content in U.S. foods is 300 mg/kg. The trialkyl and triaryl tin compounds are used as biocides and must be handled carefully.

Cost

Over the past 25 years the price of tin has varied from 50 cents/lb to its present price of abotu $4/lb. as of January 1990.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Tellurium

For vulcanization of rubber.

Atomic Number:	52
Atomic Symbol:	Te
Atomic Weight:	127.60
Electron Configuration:	$[Kr]5s^2 4d^{10} 5p^4$

History

(L. tellus, earth) Discovered by Muller von Reichenstein in 1782; named by Klaproth, who isolated it in 1798.

Sources

Tellurium is occasionally found native, but is more often found as the telluride of gold (calaverite), and combined with other metals. It is recovered commercially from the anode muds that are produced during the electrolytic refining of blister copper. The U.S., Canada, Peru, and Japan are the largest Free World producers of the element.

Properties

Crystalline tellurium has a silvery-white appearance, and when pure exhibits a metallic luster. It is brittle and easily pulverized. Amorphous tellurium is found by precipitating tellurium from a solution of telluric or tellurous acid. Whether this form is truly amorphous, or made of minute crystals, is open to question. Tellurium is a p-type semiconductor, and shows greater conductivity in certain directions, depending on alignment of the atoms.

Its conductivity increases slightly with exposure to light. It can be doped with silver, copper, gold, tin, or other elements. In air, tellurium burns with a greenish-blue flames, forming the dioxide. Molten tellurium corrodes iron, copper, and stainless steel.

Handling

Tellurium and its compounds are probably toxic and should be handled with care. Workmen exposed to as little as 0.01 mg/m^3 of air, or less, develop "tellurium breath," which has a garlic-like odor.

Isotopes

Thirty isotopes of tellurium are known, with atomic masses ranging from 108 to 137. Natural tellurium consists of eight isotopes.

Uses

Tellurium improves the machinability of copper and stainless steel, and its addition to lead decreases the corrosive action of sulfuric acid on lead and improves its strength and hardness. Tellurium is used as a basic ingredient in blasting caps, and is added to cast iron for chill control. Tellurium is used in ceramics. Bismuth telluride has been used in thermoelectric devices.

Costs

Tellurium costs about $100/lb, with a purity of about 99.5%.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Iodine

For disinfectant.

Atomic Number:	53
Atomic Symbol:	I
Atomic Weight:	126.9045
Electron Configuration:	$[Kr]5s^2 4d^{10} 5p^5$

History

(Gr. iodes, violet) Discovered by Courtois in 1811, Iodine, a halogen, occurs sparingly in the form of iodides in sea water from which it is assimilated by seaweeds, in Chilean saltpeter and nitrate-bearing earth, known as caliche in brines from old sea deposits, and in brackish waters from oil and salt wells.

Sources

Ultrapure iodine can be obtained from the reaction of potassium iodide with copper sulfate. Several other methods of isolating the element are known.

Properties

Iodine is a bluish-black, lustrous solid, volatizing at ordinary temperatures into a blue-violet gas with an irritating odor; it forms compounds with many elements, but is less active than the other halogens, which displace it from iodides. Iodine exhibits some metallic-like properties. It dissolves readily in chloroform, carbon tetrachloride, or carbon disulfide to form beautiful purple solutions. It is only slightly soluble in water.

Isotopes

Thirty isotopes are recognized. Only one stable isotope, 127I is found in nature. The artificial radioisotope 131I, with a half-life of 8 days, has been used in treating the thyroid gland. The most common compounds are the iodides of sodium and potassium (KI) and the iodates (KIO_3). Lack of iodine is the cause of goiter.

Uses

Iodine compounds are important in organic chemistry and very useful in medicine. Iodides, and thyroxine which contains iodine, are used internally in medicine, and as a solution of KI and iodine in alcohol is used for external wounds. Potassium iodide finds use in photography. The deep blue color with starch solution is characteristic of the free element.

Handling

Care should be taken in handling and using iodine, as contact with the skin can cause lesions; iodine vapor is intensely irritating to the eyes and mucus membranes. The maximum allowable concentration of iodine in air should not exceed 1 mg / m^3 (8-hour time-weighted average - 40-hour).

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 11/19/97, CST Information Services Team

Periodic Table

Xenon

For sun lamps.

Atomic Number:	54
Atomic Symbol:	Xe
Atomic Weight:	131.30
Electron Configuration:	$[Kr]5s^2 4d^{10} 5p^6$

History

(Gr. xenon, stranger) Discovered by Ramsay and Travers in 1898 in the residue left after evaporating liquid air components. Xenon is a member of the so-called noble or "inert" gases. It is present in the atmosphere to the extent of about one part in twenty million. Xenon is present in the Martian atmosphere to the extent of 0.08 ppm. the element is found in the gases evolved from certain mineral springs, and is commercially obtained by extraction from liquid air.

Isotopes

Natural xenon is composed of nine stable isotopes. In addition to these, 20 unstable isotopes have been characterized. Before 1962, it had generally been assumed that xenon and other noble gases were unable to form compounds. Evidence has been mounting in the past few years that xenon, as well as other members of zero valance elements, do form compounds. Among the "compounds" of xenon now reported are sodium perxenate, xenon deuterate, xenon hydrate, difluoride, tetrafluoride, and hexafluoride. Xenon trioxide, which is highly explosive, has been prepared. More than 80 xenon compounds have been made with xenon chemically bonded to fluorine and oxygen. Some xenon compounds are colored. Metallic xenon has been produced, using several hundred kilobars of pressure. Xenon in a vacuum tube produces a beautiful blue glow when excited by an electrical discharge.

Uses

The gas is used in making electron tubes, stoboscopic lamps, bactericidal lamps, and lamps used to excite ruby lasers for generating coherent light. Xenon is used in the nuclear energy field in bubble chambers, probes, and other applications where a high molecular weight is of value. The perxenates are used in analytical chemistry as oxidizing agents. 133Xe and 135Xe are produced by neutron irradiation in air cooled nuclear reactors. 133Xe has useful applications as a radioisotope. The element is available in sealed glass containers of gas at standard pressure. Xenon is not toxic, but its compounds are highly toxic because of their strong oxidizing

characteristics.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Barium

For spark plugs.

Atomic Number:	56
Atomic Symbol:	Ba
Atomic Weight:	137.34
Electron Configuration:	$[Xe]6s^2$

History

(Gr. barys, heavy) Baryta was distinguished from lime by Scheele in 1774; the element was discovered by Sir Humphrey Davy in 1808.

Sources

It is found only in combination with other elements, chiefly with sulfate and carbonate and is prepared by electrolysis of the chloride.

Properties

Barium is a metallic element, soft, and when pure is silvery white like lead; it belongs to the alkaline earth group, resembling calcium chemically. The metal oxidizes very easily and should be kept under petroleum or other suitable oxygen-free liquids to exclude air. It is decomposed by water or alcohol.

Uses

The metal is used as a "getter" in vacuum tubes. The most important compounds are the peroxide, chloride, sulfate, carbonate, nitrate, and chlorate. Lithopone, a pigment containing barium sulfate and zinc sulfide, has good covering power, and does not darken in the presence of sulfides. The sulfate, as permanent white is also used in paint, in X-ray diagnostic work, and in glassmaking. Barite is extensively used as a weighing agent in oil well drilling fluids, and is used in making rubber. The carbonate has been used as a rat poison, while the nitrate and chlorate give colors in pyrotechny. The impure sulfide phosphoresces after exposure to the light. All barium compounds that are water or acid soluble are poisonous. Naturally occurring barium is a mixture of seven stable isotopes. Twenty two other radioactive isotopes are known to exist.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Lanthanum

For camera lenses.

Atomic Number:	57
Atomic Symbol:	La
Atomic Weight:	138.9055
Electron Configuration:	$[Xe]6s^2 5d^1$

History

(Greek lanthanein, to lie hidden) Mosander in 1839 extracted a new earth lanthana, from impure cerium nitrate and recognized the new element.

Lanthanum was isolated in relatively pure form in 1923. Iron exchange and solvent extraction techniques have led to much easier isolation of the so-called "rare-earth" elements.

Sources

Lanthanum is found in rare-earth minerals such as cerite, monazite, allanite, and bastnasite. Monazite and bastnasite are principal ores in which lanthanum occurs in percentages up to 25 percent and 38 percent respectively. Misch metal, used in making lighter flints, contains about 25 percent lanthanum.

The availability of lanthanum and other rare earths has improved greatly in recent years. The metal can be produced by reducing the anhydrous fluoride with calcium .

Properties

Lanthanum is silvery white, malleable, ductile, and soft enough to be cut with a knife. It is one of the most reactive of the rare-earth metals. It oxidizes rapidly when exposed to air. Cold water attacks lanthanum slowly, while hot water attacks it much more rapidly.

The metal reacts directly with elemental carbon, nitrogen, boron, selenium, silicon, phosphorus, sulfur, and with halogens.

At 310C, lanthanum changes from a hexagonal to a face-centered cubic structure, and at 865C it again transforms into a body-centered cubic structure.

Isotopes

Natural lanthanum is a mixture of two stable isotopes, 138La and 139La. Twenty three other radioactive isotopes are recognized.

Uses

Rare-earth compounds containing lanthanum are extensively used in carbon lighting applications, especially by the motion picture industry for studio lighting and projection. This application consumes about 25 percent of the rare-earth compounds produced. La_2O_3 improves the alkali resistance of glass, and is used in making special optical glasses. Small amounts of lanthanum, as an additive, can be used to produce nodular cast iron.

There is current interest in hydrogen sponge alloys containing lanthanum. These alloys take up to 400 times their own volume of hydrogen gas, and the process is reversible. Every time they take up the gas, heat energy is released; therefore these alloys have possibilities in an energy conservation system.

Handling

Lanthanum and its compounds have a low to moderate acute toxicity rating; therefore, care should be taken in handling them.

Costs

The metal costs about $5/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

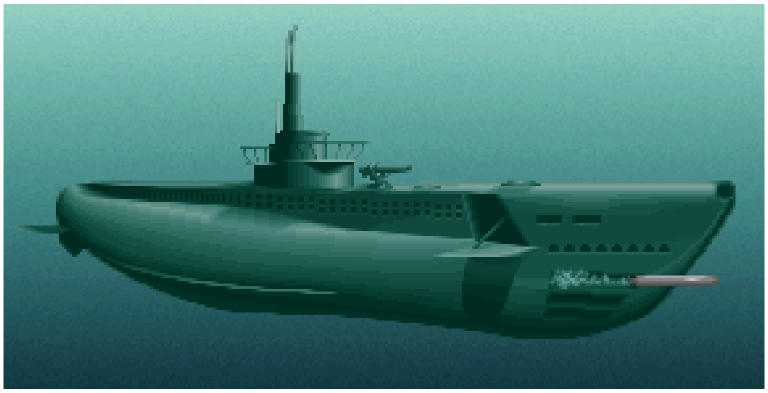
Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Hafnium

For nuclear submarines

Atomic Number:	72
Atomic Symbol:	Hf
Atomic Weight:	178.49
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^2$



History

(Hafinia, Latin name for Copenhagen) Many years before its discovery in 1932 (credited to D. Coster and G. von Hevesey), Hafnium was thought to be present in various minerals and concentrations. On the basis of the Bohr theory, the new element was expected to be associated with zirconium.

It was finally identified in zircon from Norway, by means of X-ray spectroscope analysis. It was named in honor of the city in which the discovery was made. Most zirconium minerals contain 1 to 5 percent hafnium.

It was originally separated from zirconium by repeated recrystallization of the double ammonium or potassium fluorides by von Hevesey and Jantzen. Metallic hafnium was first prepared by van Arkel and deBoer by passing the vapor of the tetraiodide over a heated tungsten filament. Almost all hafnium metal now produced is made by reducing the tetrachloride with magnesium or with sodium (Kroll Process).

Properties

Hafnium is a ductile metal with a brilliant silver luster. Its properties are considerably influenced by presence of zirconium impurities. Of all the elements, zirconium and hafnium are

two of the most difficult to separate. Although their chemistry is almost identical, the density of zirconium is about half of hafnium. Very pure hafnium has been produced, with zirconium being the major impurity.

Hafnium has been successfully alloyed with iron , titanium , niobium , tantalum , and other metals. Hafnium carbide is the most refractory binary composition known, and the nitride is the most refractory of all known metal nitrides (m.p. 3310C). At 700 degrees C hafnium rapidly absorbs hydrogen to form the composition HfH1.86.

Hafnium is resistant to concentrated alkalis, but at elevated temperatures reacts with oxygen, nitrogen, carbon , boron , sulfur , and silicon . Halogens react directly to form tetrahalides.

Uses

Because the element not only has a good absorption cross section for thermal neutrons (almost 600 times that of zirconium), but also excellent mechanical properties and is extremely corrosion-resistant, hafnium is used for reactor control rods. Such rods are used in nuclear submarines.

Hafnium is used in gas-filled and incandescent lamps, and is an efficient *getter* for scavenging oxygen and nitrogen.

Handling

Finely divided hafnium is pyrophoric and can ignite spontaneously in air. Care should be taken when machining the metal or when handling hot sponge hafnium.

Exposure to hafnium should not exceed 0.5 mg/hr. (8 hour time-weighted average - 40-hour week).

Costs

The price of the metal is in the broad range between $100/lb and $500/lb, depending on purity and quantity. The yearly demand for hafnium in the U.S. now exceeds 100,000 lb.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Tantalum

For weights.

Atomic Number:	73
Atomic Symbol:	Ta
Atomic Weight:	180.9479
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^3$

History

(Gr. Tantalos, mythological character, father of Niobe) Discovered in 1802 by Ekeberg, but many chemists thought niobium and tantalum were identical elements until Rowe in 1844, and Marignac, in 1866, showed that niobic and tantalic acids were two different acids. The early investigators only isolated the impure metal. The first relatively pure ductile tantalum was produced by von Bolton in 1903. Tantalum occurs principally in the mineral columbite-tantalite.

Sources

Tantalum ores are found in Australia, Brazil, Mozambique, Thailand, Portugal, Nigeria, Zaire, and Canada.

Production

Separation of tantalum from niobium requires several complicated steps. Several methods are used to commercially produce the element, including electrolysis of molten potassium fluorotantalate, reduction of potassium fluorotantalate with sodium, or reacting tantalum carbide with tantalum oxide. Twenty five isotopes of tantalum are known to exist. Natural tantalum contains two isotopes.

Properties

Tantalum is a gray, heavy, and very hard metal. When pure, it is ductile and can be drawn into fine wire, which is used as a filament for evaporating metals such as aluminum. Tantalum is almost completely immune to chemical attack at temperatures below 150_oC, and is attacked only by hydrofluoric acid, acidic solutions containing the fluoride ion, and free sulfur trioxide. Alkalis attack it only slowly. At high temperatures, tantalum becomes much more reactive. The element has a melting point exceeded only by tungsten and rhenium. Tantalum is used to make a variety

of alloys with desirable properties such as high melting point, high strength, good ductility, etc. Tantalum has a good "gettering" ability at high temperatures, and tantalum oxide films are stable and have good rectifying and dielectric properties.

Uses

Scientists at Los Alamos have produced a tantalum carbide graphite composite material, which is said to be one of the hardest materials ever made. The compound has a melting point of 3738$_o$C. Tantalum is used to make electrolytic capacitors and vacuum furnace parts, which account for about 60% of its use. The metal is also widely used to fabricate chemical process equipment, nuclear reactors, aircraft, and missile parts. Tantalum is completely immune to body liquids and is a nonirritating material. It has, therefore, found wide use in making surgical appliances. Tantalum oxide is used to make special glass with high index of refraction for camera lenses. The metal has many other uses.

Cost

The price of (99.9%) tantalum in Dec. 1988 was about $50 / oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Rhenium

For jewelry plating.

Atomic Number:	75
Atomic Symbol:	Re
Atomic Weight:	186.207
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^5$

History

(L. Rhenus, Rhine) Discovery of rhenium is generally attributed to Noddack, Tacke, and Berg, who announced in 1925 they had detected the element in platinum ore and columbite. They also found the element in gadolinite and molybdenite. By working up 660 kg of molybdenite in 1928 they were able to extract 1 g of rhenium.

Sources

Rhenium does not occur free in nature or as a compound in a distinct mineral species. It is, however, widely spread throughout the earth's crust to the extent of about 0.001 ppm. Commercial rhenium in the U.S. today is obtained from molybdenum roaster-flue dusts obtained from copper-sulfide ores mined in the vicinity of Miami, Arizona, and elsewhere in Arizona and Utah.

Some molybdenum contain from 0.002% to 0.2% rhenium. More than 150,000 troy ounces of rhenium are now being produced yearly in the United States. The total estimated Free World reserve of rhenium metal is 3500 tons. Rhenium metal is prepared by reducing ammonium perrhentate with hydrogen at elevated temperatures.

Isotopes

Natural rhenium is a mixture of two stable isotopes. Twenty six other unstable isotopes are recognized.

Properties

The element is silvery white with a metallic luster; its density is exceeded only by that of platinum , iridium, and osmium , and its melting point is exceeded only by that of tungsten and

carbon .

The usual commercial form of the element is powder, but it can be consolidated by pressing and resistance-sintering in a vacuum or hydrogen atmosphere. This process produces a compact shape in excess of 90 percent of the density of the metal.

Annealed rhenium is very ductile, and can be bent, coiled, or rolled. Rhenium is used as an additive to tungsten and molybdenum -based alloys to impart useful properties.

Uses

It is widely used as filaments for mass spectrographs and ion gauges. Rhenium-molybdenum alloys are superconductive at 10 K.

Rhenium is also used as an electrical contact material because it has good wear resistance and withstands arc corrosion. Thermocouples made of Re-W are used for measuring temperatures up to 2200C, and rhenium wire is used in photoflash lamps for photography.

Rhenium catalysts are exceptionally resistant to poisoning from nitrogen , sulfur , and phosphorus , and are used for hydrogenation of fine chemicals.

Costs

In 1928 rhenium cost $10,000 / g. The price today is about $250 / troy oz.

Hazards

Because little is known about its toxicity, it should be handled with care until more data becomes available.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Platinum

For dental crowns.

Atomic Number:	78
Atomic Symbol:	Pt
Atomic Weight:	195.09
Electron Configuration:	$[Xe]6s^{1}4f^{14}5d^{9}$

History

(Sp. platina, silver) Discovered in South America by Ulloa in 1735 and by Wood in 1741. The metal was used by pre-Columbian Indians.

Sources

Platinum occurs native, accompanied by small quantities of iridium, osmium, palladium, ruthenium, and rhodium, all belonging to the same group of metals. These are found in the alluvial deposits of the Ural mountains, of Columbia, and of certain western American states. Sperrylite, occurring with the nickel-bearing deposits of Sudbury, Ontario, is the source of a considerable amount of metal.

The large production of nickel makes up for the fact that is only one part of the platinum metals in two million parts of ore.

Properties

Platinum is a beautiful silvery-white metal, when pure, and is malleable and ductile. It has a coefficient of expansion almost equal to that of soda-lime-silica glass, and is therefore used to make sealed electrodes in glass systems. The metal does not oxidize in air at any temperature, but is corroded by halogens, cyanides, sulfur, and caustic alkalis.

It is insoluble in hydrochloric and nitric acid, but dissolves when they are mixed as aqua regia, forming chloroplatinic acid.

Uses

The metal is extensively used in jewelry, wire, and vessels for laboratory use, and in many valuable instruments including therocouple elements. It is also used for electrical contacts, corrosion-resistant apparatus, and in dentistry.

Platinum-cobalt alloys have magnetic properties. One such alloy made of 76.7% Pt and 23.3% Co, by weight, is an extremely powerful magnet that offers a B-H (max) almost twice that of Alnico V. Platinum resistance wires are used for constructing high-temperature electric furnaces.

The metal is used for coating missile nose cones, jet engine fuel nozzles, etc., which must perform reliably at high temperatures for long periods of time. The metal, like palladium, absorbs large volumes, of hydrogen, retaining it at ordinary temperatures but giving it up when heated.

In the finely divided state platinum is an excellent catalyst, having long been used in the contact process for producing sulfuric acid. It is also used as a catalyst in cracking petroleum products. Much interest exists in using platinum as a catalyst in fuel cells and in antipollution devices for automobiles.

Platinum anodes are extensively used in cathodic protection systems for large ships and ocean-going vessels, pipelines, steel piers, etc. Fine platinum wire will glow red hot when placed in the vapor of methyl alcohol. It acts here as a catalyst, converting the alcohol to formaldehyde. The phenomenon has been used commercially to produce cigarette lighters and hand warmers. Hydrogen and oxygen explode in the presence of platinum.

Costs

The price of platinum has varied widely. More than a century ago platinum was used to adulterate gold and was worth nearly eight times more than gold in 1920. The price in January 1990 was about $500 / troy oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Iridium

For helicopter spark plugs.

Atomic Number:	77
Atomic Symbol:	Ir
Atomic Weight:	192.22
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^7$

History

(L. iris, rainbow) Tennant discovered iridium in 1803 in the residue left when crude platinum is dissolved by aqua regia. The name *iridium* is appropriate because its salts are highly colored.

Properties

Iridium, a metal of the platinum family, is white, similar to platinum, but with a slight yellowish cast. Because iridium is very hard and brittle, it is hard to machine, form, or work.

It is the most corrosion-resistant metal known, and was used in making the standard meter bar of Paris, which is a 90 percent platinum and 10 percent iridium alloy. This meter bar was replaced in 1960 as a fundamental unit of length (see under Krypton).

Iridium is not attacked by any of the acids nor by aqua regia, but is attacked by molten salts, such as NaCl and NaCN. The specific gravity of iridium is only very slightly lower than osmium, which is generally credited as the heaviest known element. Calculations of the densities of iridium and osmium from the space lattices give values of 22.65 and 22.61 g/cm^3, respectively. These values may be more reliable than actual physical measurements. At present, therefore, we know that either iridium or osmium is the densest known element, but the data do not yet allow selection between the two.

Sources

Iridium occurs uncombined in nature with platinum and other metals of this family in alluvial deposits. It is recovered as a by-product from the nickel mining industry.

Uses

Although its principal use is as a hardening agent for platinum, iridium is also used to make crucibles and devices requiring high temperatures. It is also used for electrical contacts.

The element forms an alloy with osmium which is used for tipping pens and compass bearings.

Costs

Iridium costs about $500 / troy ounce (as of 1990).

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Osmium

For fountain pen points.

Atomic Number:	76
Atomic Symbol:	Os
Atomic Weight:	190.2
Electron Configuration:	$[Xe]6s^{2}4f^{14}5d^{6}$

History

(Gr. osme, a smell) Discovered in 1803 by Tennant in the residue left when crude platinum is dissolved by aqua regia.

Sources

Osmium occurs in iridosule and in platinum-bearing river sands of the Urals, North America, and South America. It is also found in the nickel-bearing ores of Sudbury, Ontario region along with other platinum metals. While the quantity of platinum metals in these ores is very small, the large tonnages of nickel ores processed make commercial recovery possible.

Properties

The metal is lustrous, bluish white, extremely hard, and brittle even at high temperatures. It has the highest melting point and the lowest vapor pressure of the platinum group. The metal is very difficult to fabricate, but the powdered or spongy metal slowly gives off osmium tetroxide, which as a powerful oxidizing agent and has a strong smell. The tetroxide is highly toxic, and boils at 130$_{o}$C.

The measured densities of iridium and osmium seem to indicate that osmium is slightly more dense than iridium, so osmium has generally been credited with being the heaviest known element. Calculations of the density from the space lattice which may be more reliable for these elements than actual measurements, however, give a density of 22.65 for iridium compared to 22.661 for osmium. At present, therefore, we know either iridium or osmium is the heaviest element, but the data do not allow selection between the two.

Handling

Concentrations in air as low as 10^{7} g / m^{3} can cause lung congestion, skin damage, or eye damage. Exposure to osmium tetroxide should not exceed 0.0016 mg / m^{3} (8-hour time weighted average - 40-hour work week).

Uses

The tetroxide has been used to detect fingerprints and to stain fatty tissue for miscroscope slides. The metal is almost entirely used to produce very hard alloys, with other metals of the platinum group, for fountain pen tips, instrument pivots, phonograph needles, and electrical contacts.

Cost

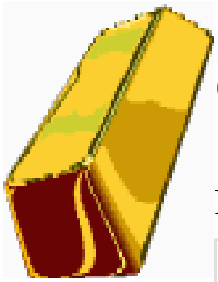
The price of 99% pure osmium powder - the form usually supplied commercially - is about $100 / g, depending on quantity and supplier.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

\

Periodic Table



Gold

For gold bars.

Atomic Number:	79
Atomic Symbol:	Au
Atomic Weight:	196.9665
Electron Configuration:	$[Xe]6s^{1}4f^{14}5d^{10}$

History

(Sanskrit Jval; Anglo-Saxon gold; L. aurum, gold) Known and highly valued from earliest times, gold is found in nature as the free metal and in tellurides; it is very widely distributed and is almost always associated with quartz or pyrite.

Sources

It occurs in veins and alluvial deposits, and is often separated from rocks and other minerals by mining and panning operations. About two thirds of the world's gold output comes from South Africa, and about two thirds of the total U.S. production comes from South Dakota and Nevada. The metal is recovered from its ores by cyaniding, amalgamating, and smelting processes. Refining is also frequently done by electrolysis. Gold occurs in sea water to the extent of 0.1 to 2 mg/ton, depending on the location where the sample is taken. As yet, no method has been found for recovering gold from sea water profitably.

Properties

It is estimated that all the gold in the world, so far refined, could be placed in a single cube 60 ft. on a side. Of all the elements, gold in its pure state is undoubtedly the most beautiful. It is metallic, having a yellow color when in a mass, but when finely divided it may be black, ruby, or purple. The Purple of Cassius is a delicate test for auric gold. It is the most malleable and ductile metal; 1 oz. of gold can be beaten out to 300 ft^2. It is a soft metal and is usually alloyed to give it more strength. It is a good conductor of heat and electricity, and is unaffected by air and most reagents.

Uses

It is used in coinage and is a standard for monetary systems in many countries. It is also extensively used for jewelry, decoration, dental work, and for plating. It is used for coating certain space satellites, as it is a good reflector of infrared and is inert.

Cost

Gold, like other precious metals, is measured in troy weight; when alloyed with other metals, the term carat is used to express the amount of gold present, 24 carats being pure gold. For many years the value of gold was set by the U.S. at $20.67/troy ounce; in 1934 this value was fixed by law at $35.00/troy ounce, 9/10th fine. On March 17, 1968, because of a gold crisis, a two-tiered pricing system was established whereby gold was still used to settle international accounts at the old $35.00/troy ounce price while the price of gold on the private market would be allowed to fluctuate. Since this time, the price of gold on the free market has fluctuated widely. The price of gold on the free market reached a price of $620/troy oz. in January 1980. As of January 1990, gold was priced at about $410/troy oz.

Isotopes

The most common gold compounds are auric chloride and chlorauric acid, the latter being used in photography for toning the silver image. Gold has 18 isotopes; 198Au, with a half-life of 2.7 days, is used for treating cancer and other diseases. Disodium aurothiomalate is administered intramuscularly as a treatment for arthritis. A mixture of one part nitric acid with three of hydrochloric acid is called aqua regia (because it dissolved gold, the King of Metals). Gold is available commercially with a purity of 99.999+%. For many years the temperature assigned to the freezing point of gold has been 1063.0C; this has served as a calibration point for the International Temperature Scales (ITS-27 and ITS-48) and the International Practical Temperature Scale (IPTS-48). In 1968, a new International Practical Temperature Scale (IPTS-68) was adopted, which demands that the freezing point of gold be changed to 1064.43C. The specific gravity of gold has been found to vary considerably depending on temperature, how the metal is precipitated, and cold-worked.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Thallium

For insecticides.

Atomic Number:	81
Atomic Symbol:	Tl
Atomic Weight:	204.37
Electron Configuration:	[Xe]$6s^2 4f^{14} 5d^{10} 6p^1$

History

(Gr. thallos, a green shoot or twig) Thallium was discovered spectroscopically in 1861 by Crookes. The element was named after the beautiful green spectral line, which identified the element. The metal was isolated both by Crookes and Lamy in 1862 about the same time.

Sources

Thallium occurs in crooksite, lorandite, and hutchinsonite. It is also present in pyrites and is recovered from the roasting of this ore in connection with the production of sulfuric acid. It is also obtained from the smelting of lead and zinc ores. Extraction is somewhat complex and depends on the source of the thallium. Manganese nodules, found on the ocean floor, contain thallium.

Properties

When freshly exposed to air, thallium exhibits a metallic luster, but soon develops a bluish-gray tinge, resembling lead in appearance. A heavy oxide builds up on thallium if left in air, and in the presence of water the hydride is formed. The metal is very soft and malleable. It can be cut with a knife. Twenty five isotopic forms of thallium, with atomic masses ranging from 184 to 210 are recognized. Natural thallium is a mixture of two isotopes. A mercury-thallium alloy, which forms a eutectic at 8.5% thallium, is reported to freeze at -60C, some 20 degrees below the freezing point of mercury.

Cost

Commercial thallium metal (99%) costs about $40/lb.

Handling

The element and its compounds are toxic and should be handled carefully. Contact of the metal with skin is dangerous, and when melting the metal adequate ventilation should be provided. Exposure to thallium (soluble compounds) - skin, as Tl, should not exceed 0.1 mg / m^3 (8-hour time-weighted average - 40-hour work week). Thallium is suspected of carcinogenic potential for man.

Uses

Thallium sulfate has been widely employed as a rodenticide and ant killer. It is odorless and tasteless, giving no warning of its presence. Its use, however, has been prohibited in the U.S. since 1975 as a household insecticide and rodenticide. The electrical conductivity of thallium sulfide changes with exposure to infrared light, and this compound is used in photocells. Thallium bromide-iodide crystals have been used as infrared optical materials. Thallium has been used, with sulfur or selenium and arsenic, to produce low melting glasses with become fluid between 125 and 150C. These glasses have properties at room temperatures similar to ordinary glasses and are said to be durable and insoluble in water. Thallium oxide has been used to produce glasses with a high index of refraction. Thallium has been used in treating ringworm and other skin infections; however, its use has been limited because of the narrow margin between toxicity and therapeutic benefits.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Bismuth

For fuses.

Atomic Number:	83
Atomic Symbol:	Bi
Atomic Weight:	208.9804
Electron Configuration:	$[Xe]6s^{2}4f^{14}5d^{10}6p^{3}$

History

(Ger. Weisse Masse, white mass; later Wisuth and Bisemutum) In early times bismuth was confused with tin and lead. Claude Geoffroy the Younger showed it to be distinct from lead in 1753.

Properties

It is a white crystalline, brittle metal with a pinkish tinge. It occurs native. Bismuth is the most diamagnetic of all metals, and the thermal conductivity is lower than any metal, except mercury. It has a high electrical resistance, and has the highest Hall effect of any metal (i.e., greatest increase in electrical resistance when placed in a magnetic field).

Sources

The most important ores are bismuthinite or bismuth glance and bismite. Peru, Japan, Mexico, Bolivia, and Canada are major bismuth producers. Much of the bismuth produced in the U.S. is obtained as a by-product in refining lead, copper, tin, silver, and gold ores.

Uses

"Bismanol" is a permanent magnet of high coercive force, made of MnBi, by the U.S. Naval Surface Weapons Center. Bismuth expands 3.32% on solidification. This property makes bismuth alloys particularly suited to the making of sharp castings of objects subject to damage by high temperatures. With other metals such as tin, cadmium, etc., bismuth forms low-melting alloys which are extensively used for safety devices in fire detection and extinguishing systems. Bismuth is used in producing malleable irons and is finding use as a catalyst for making acrylic fibers. When bismuth is heated in air it burns with a blue flame, forming yellow fumes of the oxide. The metal is also used as a thermocouple material, and has found application as a carrier

for U235 or U233 fuel in nuclear reactors. Its soluble salts are characterized by forming unsoluble basic salts on the addition of water, a property sometimes used in detection work. Bismuth oxychloride is used extensively in cosmetics. Bismuth subnitrate and subcarbonate are used in medicine.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Polonium

For nuclear batteries.

Atomic Number:	84
Atomic Symbol:	Po
Atomic Weight:	209
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^{10} 6p^4$

History

(Poland, native country of Mme. Curie) Polonium, also called Radium F, was the first element discovered by Mme. Curie in 1898 while seeking the cause of radioactivity of pitchblend from Joachimsthal, Bohemia. The electroscope showed it separating with bismuth.

Sources

Polonium is a very rare natural element. Uranium ores contain only about 100 micrograms of the element per ton. Its abundance is only about 0.2% of that of radium.

In 1934, scientists discovered that when they bombarded natural bismuth (209Bi) with neutrons, 210Bi, the parent of polonium, was obtained. Milligram amounts of polonium may now be prepared this way, by using the high neutron fluxes of nuclear reactors.

Polonium is available commercially on special order from the Oak Ridge National Laboratory.

Properties

Polonium-210 is a low-melting, fairly volatile metal, 50% of which is vaporized in air in 45 hours at 55C. It is an alpha emitter with a half-life of 138.39 days. A milligram emits as many alpha particles as 5 g of radium.

The energy released by its decay is so large (140W/g) that a capsule containing about half a gram reaches a temperature above 500C. The capsule also presents a contact gamma-ray dose rate of 0.012 Gy/h. A few curies (1 curie = $3.7 \times 10_{10}$Bq) of polonium exhibit a blue glow, caused by excitation of the surrounding gas.

Polonium is readily dissolved in dilute acids, but is only slightly soluble in alkali. Polonium slats of organic acids char rapidly; halide amines are reduced to the metal.

Uses

Because almost all alpha radiation is stopped within the solid source and its container, giving up its energy, polonium has attracted attention for uses as a lightweight heat source for thermoelectric power in space satellites.

Polonium can be mixed or alloyed with beryllium to provide a source of neutrons. The element has been used in devices for eliminating static charges in textile mills, etc.; however, beta sources are both more commonly used and less dangerous. It is also used on brushes for removing dust from photographic films. The polonium for these is carefully sealed and controlled, minimizing hazards to the user.

Isotopes

Twenty five isotopes of polonium are known, with atomic masses ranging from 194 to 218. Polonium-210 is the most readily available. Isotopes of mass 209 (half-life 103 years) and mass 208 (half-life 2.9 years) can be prepared by alpha, proton, or deuteron bombardment of lead or bismuth in a cyclotron, but these are expensive to produce.

Metallic polonium has been prepared from polonium hydroxide and some other polonium compounds in the presence of concentrated aqueous or anhydrous liquid ammonia. Two allotropic modifications are known to exist.

Handling

Polonium-210 is very dangerous to handle in even milligram or microgram amounts, and special equipment and strict control is necessary. Damage arises from the complete absorption of the energy of the alpha particle into tissue.

The maximum permissible body burden for ingested polonium is only 0.03 microcuries, which represents a particle weighing only 6.8×10^{-12} g. Weight for weight it is about 2.5×10^{11} times as toxic as hydrocyanic acid. The maximum allowable concentration for soluble polonium compounds in air is about 2×10^{-11} microcuries / cm^3.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Astatine

Seldom found in nature.

Atomic Number:	85
Atomic Symbol:	At
Atomic Weight:	210
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^{10} 6p^5$

History

(Gr. astatos, unstable) Synthesized in 1940 by D.R. Corson, K.R. MacKenzie, and E. Segre at the University of California by bombarding bismuth with alpha particles. The longest-lived isotopes, with naturally occurring uranium and thorium isotopes, and traces of 217At are equilibrium with 233U and 239Np resulting from integration of thorium and uranium with naturally produced neutrons. The total amount of astatine present in the earth's crust, however, is less than 1 oz.

Uses

Krypton-85 has been used for over 25 years to measure the density of paper as it is amanufactured. The total weight of paper can be controlled to a very accurate degree by the use of krypton 85 and other radioactive nuclides. The common name for such a device is a beta gague that can measure the thickness of a material.

Production

Astatine can be produced by bombarding bismuth with energetic alpha particles to obtain the relatively long-lived 209-211At, which can be distilled from the target by heating in air.

Properties

The "time of flight" mass spectrometer has been used to confirm that this highly radioactive halogen behaves chemically very much like other halogens, particularly iodine. Astatine is said to be more metallic than iodine, and, like iodine, it probably accumulates in the thyroid gland. Workers at the Brookhaven National Laboratory have recently used reactive scattering in crossed molecular beams to identify and measure elementary reactions involving astatine.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Radon

For earthquake prediction.

Atomic Number:	86
Atomic Symbol:	Rn
Atomic Weight:	222
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^{10} 6p^6$

History

(From radium; called niton at first, L. nitens, shining) The element was discovered in 1900 by Dorn, who called it radium emanation. In 1908 Ramsay and Gray, who named it niton, isolated the element and determined its density, finding it to be the heaviest known gas. It is essentially inert and occupies the last place in the zero group of gases in the Periodic Table. Since 1923, it has been called radon.

Isotopes

Twenty isotopes are known. Radon-22, from radium, has a half-life of 3.823 days and is an alpha emitter; Radon-220, emanating naturally from thorium and called thoron, has a half-life of 55.6 s and is also an alpha emitter. Radon-219 emanates from actinium and is called actinon. It has a half-life of 3.96 s and is also an alpha emitter. It is estimated that every square mile of soil to a depth of 6 inches contains about 1 g of radium, which releases radon in tiny amounts into the atmosphere. Radon is present in some spring waters, such as those at Hot Springs, Arkansas.

Properties

On the average, one part of radon is present ot $1 \times 10_{21}$ part of air. At ordinary temperatures radon is a colorless gas; when cooled below the freezing point, radon exhibits a brilliant phosphorescence which becomes yellow as the temperature is lowered and orange-red at the temperature of liquid air. It has been reported that fluorine reacts with radon, forming a fluoride. Radon clathrates have also been reported.

Uses

Radon is still produced for therapeutic use by a few hospitals by pumping it from a radium source and sealing it in minute tubes, called seeds or needles, for application to patient. This

practice has been largely discontinued as hospitals can get the seeds directly from suppliers, who make up the seeds with the desired activity for the day of use.

Cost

Radon is available at a cost of about $4/m.

Handling

Care must be taken in handling radon, as with other radioactive materials. The main hazard is from inhalation of the element and its solid daughters which are collected on dust in the air. Good ventilation should be provided where radium, thorium, or actinium is stored to prevent build-up of the element. Radon build-up is a health consideration in uranium mines. Recently radon build-up in homes has been a concern. Many deaths from lung cancer are caused by radon exposure. In the U.S. it is recommended that remedial action be taken if the air in homes exceeds 4 pCi/l.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Francium

Seldom found in nature.

Atomic Number:	87
Atomic Symbol:	Fr
Atomic Weight:	223
Electron Configuration:	[Rn]$7s^1$

History

(France) Discovered in 1939 by Mlle. Marguerite Perey of the Curie Institute, Paris. Francium, the heaviest known member of the alkali metals series, occurs as a result of an alpha disintegration of actinium. It can also be made by artificially bombarding thorium with protons. While it occurs naturally in uranium minerals, there is probably less than an ounce of francium at any time in the total crust of the earth. It has the highest equivalent weight of any element, and is the most unstable of the first 101 elements of the periodic system. Thirty-three isotopes of francium are recognized. The longest lived 223Fr (Ac, K), a daughter of 227Ac, has a half-life of 22 min. This is the only isotope of francium occurring in nature. Because all known isotopes of francium are highly unstable, knowledge of the chemical properties of this element comes from radiochemical techniques. No weighable quantity of the element has been prepared or isolated. The chemical properties of francium most resemble cesium.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Radium

For glow in the dark paint.

Atomic Number:	88
Atomic Symbol:	Ra
Atomic Weight:	226.0254
Electron Configuration:	$[Rn]7s^2$

History

(L. radius, ray) Radium was discovered in 1898 by Mme. Curie in the pitchblende or uraninite of North Bohemia, where it occurs. There is about 1 g of radium in 7 tons of pitchblende. The element was isolated in 1911 by Mme. Curie and Debierne by; the electrolysis of a solution of pure radium chloride, employing a mercury cathode; on distillation in an atmosphere of hydrogen this amalgam yielded the pure metal.

Sources

Originally, radium was obtained from the rich pitchblende ore found in Joachimsthal, Bohemia. The carnotite sands of Colorado furnish some radium, but richer ores are found in the Republic of Zaire and the Great Lake region of Canada. Radium is present in all uranium minerals, and could be extracted, if desired, from the extensive wastes of uranium processing. Large uranium deposits are located in Ontario, New Mexico, Utah, Australia, and elsewhere.

Properties

Radium is obtained commercially as bromide and chloride; it is doubtful if any appreciable stock of the isolated element now exists. The pure metal is brilliant white when freshly prepared, but blackens on exposure to air, probably due to formation of the nitride. It exhibits luminescence, as do its slats; it decomposes in water and is somewhat more volatile than barium. It is a member of the alkaline-earth group of metals. Radium imparts a carmine red color to a flame. Radium emits alpha, beta, and gamma rays and when mixed with beryllium produce neutrons. One gram of 226Ra undergoes 3.7×10^{10} disintegrations per s. The curie is defined as that amount of radioactivity which has the same disintegration rate as 1 g of 226Ra. Twenty five isotopes are now known; radium 226, the common isotope, has a half-life of 1600 years.

Uses

One gram of radium produces about 0.0001 ml (stp) of emanation, or radon gas, per day. This is purged from the radium and sealed in minute tubes, which are used in the treatment of cancer and other diseases. Radium is used in the producing of self-luminous paints, neutron sources, and in medicine for the treatment of disease. Some of the more recently discovered radioisotopes, such as 60Co, are now being used in place of radium. Some of these sources are much more powerful, and others are safer to use. Radium loses about 1% of its activity in 25 years, being transformed into elements of lower atomic weight. Lead is a final product of disintegration. Stored radium should be ventilated to prevent build-up of radon.

Handling

Inhalation, injection, or body exposure to radium can cause cancer and other body disorders. The maximum permissible border in the total body for 226Ra is 7400 becquerel.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Actinium

Neutron source.

Atomic Number:	89
Atomic Symbol:	Ac
Atomic Weight:	227
Electron Configuration:	$[Rn]7s^2 6d^1$

History

(Gr. aktis, aktinos, beam or ray). Discovered by Andre Debierne in 1899 and independently by F. Giesel in 1902. Occurs naturally in association with uranium minerals. Actinium-227, a decay product of uranium-235, is a beta emitter with a 21.6-year half-life. Its principal decay products are thorium-227 (18.5-day half-life), radium-223 (11.4-day half-life), and a number of short-lived products including radon, bismuth, polonium, and lead isotopes. In equilibrium with its decay products, it is a powerful source of alpha rays. Actinium metal has been prepared by the reduction of actinium fluoride with lithium vapor at about 1100 to 1300-degrees C. The chemical behavior of actinium is similar to that of the rare earths, particularly lanthanum. Purified actinium comes into equilibrium with its decay products at the end of 185 days, and then decays according to its 21.6-year half-life. It is about 150 times as active as radium, making it of value in the production of neutrons.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Rutherfordium

Proposed name

Atomic Number:	104
Atomic Symbol:	Rf / Ku
Atomic Weight:	261
Electron Configuration:	$[Rn]7s^2 5f^{14} 6d^2$

History

In 1964, workers at the Joint Nuclear Research Institute at Dubna (U.S.S.R.) bombarded plutonium with accelerated 113 to 115 MeV neon ions. By measuring fission tracks in a special glass with a microscope, they detected an isotope that decays by spontaneous fission. They suggested that this isotope, which had a half-life of 0.3 +/- 0.1 s might be 260-104, produced by the following reaction: 242Pu + 22Ne --> 104 +4n.

Element 104, the first transactinide element, is expected to have chemical properties similar to those of hafnium. It would, for example, form a relatively volatile compound with chlorine (a tetrachloride).

The Soviet scientists have performed experiments aimed at chemical identification, and have attempted to show that the 0.3-s activity is more volatile than that of the relatively nonvolatile actinide trichlorides. This experiment does not fulfill the test of chemically separating the new element from all others, but it provides important evidence for evaluation.

New data, reportedly issued by Soviet scientists, have reduced the half-life of the isotope they worked with from 0.3 to 0.15 s. The Dubna scientists suggest the name *kurchatovium* and symbol Ku for element 104, in honor of Igor Vasilevich Kurchatov (1903-1960), former Head of Soviet Nuclear Research.

Isotopes

In 1969 Ghiorso, Nurmia, Harris, K.A.Y. Eskola, and P.L. Eskola of the University of California at Berkeley reported that they had positively identified two, and possibly three, isotopes of Element 104. The group indicated that, after repeated attempts, they produced isotope 260-104 reported by the Dubna groups in 1964.

The discoveries at Berkeley were made by bombarding a target of 249Cf with 12C nuclei of 71 MeV, and 13C nuclei of 69 MeV. The combination of 12C with 249Cf followed by instant emission of four neutrons produced Element 257-104. This isotope has a half-life of 4 to 5 s,

decaying by emitting an alpha particle into 253No, with a half-life of 105 s.

The same reaction, except with the emission of three neutrons, was thought to have produced 258-104 with a half-life of about 1/100 s.

Element 259-104 is formed by the merging of a 13C nuclei with 249Cf, followed by emission of three neutrons. This isotope has a half-life of 3 to 4 s, and decays by emitting an alpha particle into 255No, which has a half-life of 185 s.

Thousands of atoms of 257-104 and 259-104 have ben detected. The Berkeley group believes their identification of 258-104 is correct, but attaches less confidence to this work than to their work on 257-104 and 259-104.

The claims for discovery and the naming of Element 104 are still in question. The Berkeley group proposes for the new element the name *rutherfordium* (symbol Rf), in honor of Ernest R. Rutherford, a New Zealand physicist. Meanwhile, the International Union of Pure and Applied Physics has proposed using the neutral temporary name, *unnilquadium*.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Dubnium

Proposed Name

Atomic Number:	105
Atomic Symbol:	Db
Atomic Weight:	262
Electron Configuration:	$[Rn]7s^{2}5f^{14}6d^{3}$

History

In 1967 G.N. Flerov reported that a Soviet team working at the Joint Institute for Nuclear Research at Dubna may have produced a few atoms of 260-105 and 261-105 by bombarding 243Am with 22Ne. The evidence was based on time-coincidence measurements of alpha energies.

In 1970 Dubna scientists synthesized Element 105 and, by the end of April 1970, "had investigated all the types of decay of the new element and had determined its chemical properties," according to a report in 1970. The Soviet group had not proposed a name for 105. In late April 1970, it was announced that Ghiorso, Nurmia, Haris, K.A.Y. Eskola, and P.L. Eskola, working at the University of California at Berkeley, had positively identified element 105. The discovery was made by bombarding a target of 249Cf with a beam of 84 MeV nitrogen nuclei in the Heavy Ion Linear Accelerator (HILAC). When a 15N nuclear is absorbed by a 249Cf nucleus, four neutrons are emitted and a new atom of 260-105 with a half-life of 1.6 s is formed. While the first atoms of Element 105 are said to have been detected conclusively on March 5, 1970, there is evidence that Element 105 had been formed in Berkeley experiments a year earlier by the method described.

Ghiorso and his associates have attempted to confirm Soviet findings by more sophisticated methods without success. The Berkeley Group proposed the name *hahnium* -- after the late German scientist Otto Hahn (1879-1968) -- and symbol Ha. However, the International Union of Pure and Applied Chemistry panel members in 1977 recommended that element 105 be named to Dubnium (symbol Db) after the site of the Joint Institute for Nuclear Research in Russia. Unfortunately, the name hahnium will not be used again according to the rules for naming new elements. Some scientists still use the earlier name of hahnium because it had been used for about 25 years.

Isotopes

In October 1971, it was announced that two new isotopes of element 105 were synthesized with the heavy ion linear accelerator by A. Ghiorso and co-workers a Berkeley. Element 261-105 was produced both by bombarding 250Cf with 15N and by bombarding 249Bk with 16O. The isotope emits 8.93-MeV alpha particles and decays to 257Lr with a half-life of about 1.8 s. Element 262-105 was produced by bombarding 249Bk with 18O. It emits 8.45 MeV alpha particles and decays to 258Lr with a half-life of about 40 s. Seven isotopes of element 105 (unnilpentium) are now recognized.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Sg Seaborgium

Proposed Name

Atomic Number:	106
Atomic Symbol:	Sg
Atomic Weight:	263
Electron Configuration:	$[Rn]7s^{2}5f^{14}6d^{4}$

History

In June 1974, members of the Joint Institute for Nuclear Research in Dubna, U.S.S.R., reported their discovery of Element 106, which they reported to have synthesized. Glenn Seaborg was part of this group, and the element was named in his honor. Seaborgium is often still referred to as Element 106 because the international committee in charge of names changed the rules. They decided retroactively it couldn't be named after a living person.

In September 1974, workers of the Lawrence Berkeley and Livermore Laboratories also claimed creation Element 106 "without any scientific doubt." The LBL and LLL Group used the Super HILAC to accelerate 18O ions onto a 249Cf target.

Element 106 was created by the reaction 249Cf(18O, 4N)263X, which decayed by alpha emission to rutherfordium, and then by alpha emission to nobelium, which in turn further decayed by alpha between daughter and granddaughter. The element so identified had alpha energies of 9.06 and 9.25 MeV with a half-life of 0.9 +/- 0.2 s.

At Dubna, 280-MeV ions of 54Cr from the 310-cm cyclotron were used to strike targets of 206Pb, 207Pb, and 208Pb, in separate runs. Foils exposed to a rotating target disc were used to detect spontaneous fission activities. The foils were etched and examined microscopically to detect the number of fission tracks and the half-life of the fission activity.

Other experiments were made to aid in confirmation of the discovery. Neither the Dubna team nor the Berkeley-Livermore Group has proposed a name as of yet for element 106 (unnilhexium).

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Nobelium

Atomic Number:	102
Atomic Symbol:	No
Atomic Weight:	259
Electron Configuration:	$[Rn]7s^2 5f^{14}$

History

(Alfred Nobel, discoverer of dynamite) Nobelium was unambiguiously discovered and identified in April 1958 at Berkeley by A. Ghiorso, T. Sikkeland, J.R. Walton, and G.T. Seaborg, who used a new double-recoil technique. A heavy-ion linear accelerator (HILAC) was used to bombard a thin target of curium (95% 244Cm and 4.5% 246Cm) with 12C ions to produce 102No according to the 246Cm(12C, 4n) reaction.

In 1957 workers in the United States, Britain, and Sweden announced the discovery of an isotope of element 102 with a 10-minute half-life at 8.5 MeV, as a result of bombarding 244Cm with 13C nuclei. On the basis of this experiment, the name *nobelium* was assigned and accepted by the Commission on Atomic Weights of the International Union of Pure and Applied Chemistry.

The acceptance of the name was premature because both Russian and American efforts now completely rule out the possibility of any isotope of Element 102 having a half-life of 10 min in the vicinity of 8.5 MeV. Early work in 1957 on the search for this element, in Russia at the Kurchatov Institute, was marred by the assignment of 8.9 +/- 0.4 MeV alpha radiation with a half-life of 2 to 40 sec, which was too indefinite to support discovery claims.

Confirmatory experiments at Berkeley in 1966 have shown the existence of 254-102 with a 55-s half-life, 252-102 with a 2.3-s half-life, and 257-102 with a 23-s half-life.

Following tradition giving the right to name an element to the discoverer(s), the Berkeley group in 1967, suggested that the hastily given name *nobelium* along with the symbol *No* , be retained.

Isotopes

Ten isotopes are now recognized, one of which -- 255-102 -- has a half-life of 3 minutes.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Bh Bohrium

Formally known as Ns Nielsbohrium

Proposed Name

Atomic Number:	107
Atomic Symbol:	
Atomic Weight:	262
Electron Configuration:	$[Rn]7s^2 5f^{14} 6d^5$

History

In 1976 Soviet scientists at Dubna announced they had synthesized element 107 by bombarding 204Bi with heavy nuclei of 54Cr. Reports say that experiments in 1975 had allowed scientists "to glimpse" the new element for 2/1000 s. A rapidly rotating cylinder, coated with a thin layer of bismuth metal, was used as a target. This was bombarded by a stream of 54Cr ions fired tangentially.

The existence of element 107 was confirmed by a team of West German physicists at the Heavy Ion Research Laboratory at Darmstadt, who created and identified six nuclei of element 107.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 10/20/1999, CST Information Services Team

Periodic Table

Hassium

Proposed Name

Atomic Number:	108
Atomic Symbol:	Hs
Atomic Weight:	265
Electron Configuration:	$[Rn]7s^2 5f^{14} 6d^6$

History

(origin is Latin "Hassias" meaning "Hess", the German state.) Discovered by Peter Armbruster, Gottfried Munzenber and co-workers at GSI in Darmstadt, Germany in 1984.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Meitnerium

Proposed Name

Atomic Number:	109
Atomic Symbol:	Mt
Atomic Weight:	265
Electron Configuration:	$[Rn]7s^2 5f^{14} 6d^7$

History

On August 29, 1982, physicists at the Heavy Ion Research Laboratory, Darmstadt, West Germany made and identified element 109 by bombing a target of Bi-209 with accelerated nuclei of Fe-58. If the combined energy of two nuclei is sufficiently high, the repulsive forces between the nuclei can be overcome.

In this experiment a week of target bombardment was required to produce a single fused nucleus. The team confirmed the existence of element 109 by four independent measurements. The newly formed atom recoiled from the target at predicted velocity and was separated from smaller, faster nuclei by a newly developed velocity filter. The time of flight to the detector and the striking energy were measured and found to match predicted values.

The nucleus of 266X started to decay 5 ms after striking the detector. A high-energy alpha particle was emitted, producing 267/107X. This in turn emitted an alpha particle, becoming 258/105Ha, which in turn captured an electron and became 258/104Rf. This in turn decayed into other nuclides. This experiment demonstrated the feasibility of using fusion techniques as a method of making new, heavy nuclei.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Element 110

Unnamed

Atomic Number:	110
Atomic Symbol:	
Atomic Weight:	
Electron Configuration:	

Press Release:Discovery of Element 110 at GSI

November 9, 1994 at 4:39 pm, the first atom of the heaviest chemical atom with atomic number 110 was detected at the Gesellschaft fur Schwerionenforschung (GSI) in Darmstadt, in Germany. For the last ten years, this element has been the subject of an intense search by many laboratories world-wide.

Discovery

Element 110 was produced by fusing nickel and lead atom together. This was achieved by accelerating the nickel atoms to a high energy in the heavy ion accelerator. "This rare reaction occurs only at a very specific velocity of the nickel projectile. Over a period of many days, many billion billion nickel atoms must be shot at a lead target in order to produce and identify a single atom of element 110. The atoms produced in the nickel-lead collisions are selected by a velocity filter and then captured in a detector system which measures their decay. The energy of the emitted helium nuclei serves to identify the atom" (Press Release). This element was only found to have a lifetime of less than 1/1000th of a second. It is expected that soon a heavier version of element 110 that might be more stable, and that lives slightly longer will be developed.

Sources: Element 110 Discovered and the GSI.

Last Updated: 12/19/97 CST Information Services Team

Periodic Table

Element 111

Atomic Number:	111
Atomic Symbol:	
Atomic Weight:	
Electron Configuration:	

History

Discovered by Gesellschaft Schwerionenforschung (GSI) in Darmstadt, in 1994. Reasearch group of S. Hofmann, V. Ninov, F.P. Hessberger, P. Armbruster, H. Folger, G. Munzenberg, H.J. Schott, and others.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Element 112

Atomic Number:	112
Atomic Symbol:	
Atomic Weight:	277
Electron Configuration:	

History

February 9, 1996 at 10:37 pm, at the Gesellschaft fur Schwerionenforschung in Darmstadt, Germany a team of scientists discovered their sixth element. This element has the atomic number 112 and is currently the heaviest element ever produced by man. It has an atomic mass of 277.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

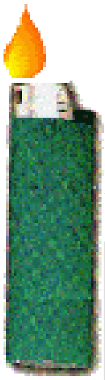
Last Updated: 12/19/97, CST Information Services Team

Information on elements 114, 116, and 118 is not yet available.

These elements have, however, been produced and can exist.

Expect an update cerca Y2K

Periodic Table



Cerium

For lighter flints.

Atomic Number:	58
Atomic Symbol:	Ce
Atomic Weight:	140.12
Electron Configuration:	$[Xe]6s^2 4f^1 5d^1$

History

Cerium was named for the asteroid Ceres, which was discovered in 1801. The element was discovered two years later in 1803 by Klaproth and by Berzelius and Hisinger. In 1875 Hillebrand and Norton prepared the metal.

Sources

Cerium is the most abundant so-called rare-earths metal. It is found in a number of minerals including allanite (also known as orthite), monazite, bastnasite, certie, and samarskite. Monazite and bastnasite are presently the two more important sources of cerium.

Large deposits of monazite (found on the beaches of Travancore, India and in river sands in Brazil), allanite (in the western United States), and bastnasite (in Southern California) will supply cerium, thorium, and the other rare-earth metals for many years to come.

Metallic cerium is prepared by metallothermic reduction techniques, such as reducing cerous fluoride with calcium, or using electrolysis of molten cerous chloride or others processes. The metallothermic technique produces high-purity cerium.

Properties

Cerium is especially interesting because of its variable electronic structure. The energy of the inner 4f level is nearly the same as that of the outer or valence electrons, and only small amounts of energy are required to change the relative occupancy of these electronic levels. This gives rise to dual valency states.

For example, a volume change of about 10 percent occurs when cerium is subjected to high pressures or low temperatures. Cesium's valence appears to change from about 3 to 4 when it is cooled or compressed. The low temperature behavior of cerium is complex.

Cerium is an iron-gray lustrous metal. It is malleable, and oxidizes very readily at room temperature, especially in moist air. Except for europium, cerium is the most reactive of the *rare-earth* metals. It decomposes slowly in cold water and rapidly in hot water.

Alkali solutions and dilute and concentrated acids attack the metal rapidly. The pure metal is likely to ignite if scratched with a knife.

Ceric slats are orange red or yellowish; cerous salts are usually white.

Uses

Cerium is a component of misch metal, which is extensively used in the manufacture of pyrophoric alloys for cigarette lighters. While cerium is not radioactive, the impure commercial grade may contain traces of thorium, which is radioactive. The oxide is an important constituent of incandescent gas mantles and is emerging as a hydrocarbon catalyst in *self cleaning* ovens. In this application it can be incorporated into oven walls to prevent the collection of cooking residues.

As ceric sulfate is used extensively as a volumetric oxidizing agent in quantitative analysis. Cerium compounds are used in the manufacture of glass, both as a component and as a decolorizer.

The oxide is finding increased use as a glass polishing agent instead of rouge, for it polishes much faster than rouge. Cerium, with other rare earths, is used in carbon-arc lighting, especially in the motion picture industry. It is also finding use as an important catalyst in petroleum refining and in metallurgical and nuclear applications.

Costs (estimated)

In small lots, 99.9% cerium costs about $125/kg.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Thorium

For crucibles.

Atomic Number:	90
Atomic Symbol:	Th
Atomic Weight:	232.0381
Electron Configuration:	$[Rn]7s^2 6d^2$

History

(Thor, Scandinavian god of war) Discovered by Berzelius in 1828. Much of the internal heat the earth produces has been attributed to thorium and uranium. Because of its atomic weight, valence, etc., it is now considered to be the second member of the actinide series of elements.

Sources

Thorium occurs in thorite and in thorianite. Large deposits of thorium minerals have been reported in New England and elsewhere, but these have not yet been exploited. Thorium is now thought to be about three times as abundant as uranium and about as abundant as lead or molybdenum. Thorium is recovered commercially from the mineral monazite, which contains from 3 to 9% ThO_2 along with rare-earth minerals.

Uses

The metal is a source of nuclear power. There is probably more energy available for use from thorium in the minerals of the earth's crust than from both uranium and fossil fuels. Any sizable demand from thorium as a nuclear fuel is still several years in the future. Work has been done in developing thorium cycle converter-reactor systems. Several prototypes, including the HTGR (high-temperature gas-cooled reactor) and MSRE (molten salt converter reactor experiment), have operated. While the HTGR reactors are efficient, they are not expected to become important commercially for many years because of certain operating difficulties.

Production

Several methods are available for producing thorium metal; it can be obtained by reducing thorium oxide with calcium, by electrolysis of anhydrous thorium chloride in a fused mixture of sodium and potassium chlorides, by calcium reduction of thorium tetrachloride mixed with

anhydrous zinc chloride, and by reduction of thorium tetrachloride with an alkali metal. Thorium was originally assigned a position in Group IV of the periodic table.

Properties

When pure, thorium is a silvery-white metal which is air-stable and retains its luster for several months. When contaminated with the oxide, thorium slowly tarnishes in air, becoming gray and finally black. The physical properties of thorium are greatly influenced by the degree of contamination with the oxide. The purest specimens often contain several tenths of a percent of the oxide. High-purity thorium has been made. Pure thorium is soft, very ductile, and can be cold-rolled, swaged, and drawn. Thorium is dimorphic, changing at 1400C from a cubic to a body-centered cubic structure. Thorium oxide has a melting point of 3300C, which is the highest of all oxides. Only a few elements, such as tungsten, and a few compounds, such as tantalum carbide, have higher melting points. Thorium is slowly attacked by water, but does not dissolve readily in most common acids, except hydrochloric. Powdered thorium metal is often pyrophoric and should be handled carefully. When heated in air, thorium turnings ignite and burn brilliantly with a white light.

Uses

The principal use of thorium has been in the preparation of the Welsbach mantle, used for portable gas lights. These mantles, consisting of thorium oxide with about 1% cerium oxide and other ingredients, glow with a dazzling light when heated in a gas flame. Thorium is an important alloying element in magnesium, imparting high strength and creep resistance at elevated temperatures. Because thorium has a low work-function and high electron emission, it is used to coat tungsten wire used in electronic equipment. The oxide is also used to control the grain size of tungsten used for electric lamps; it is also used for high-temperature laboratory crucibles. Glasses containing thorium oxide have a high refractive index and low dispersion. Consequently, they find application in high quality lenses for cameras and scientific instruments. Thorium oxide has also found use as a catalyst in the conversion of ammonia to nitric acid, in petroleum cracking, and in producing sulfuric acid.

Isotopes

Twenty five isotopes of thorium are known with atomic masses ranging from 212 to 236. All are unstable. 232Th occurs naturally and has a half-life of 1.4 x 10^{10} years. It is an alpha emitter. 232Th goes through six alpha and four beta decay steps before becoming the stable isotope 208Pb. 232Th is sufficiently radioactive to expose a photographic plate in a few hours. Thorium disintegrates with the production of "thoron" (220Rn), which is an alpha emitter and presents a radiation hazard. Good ventilation of areas where thorium is stored or handled is therefore essential.

Cost

Thorium metal (99.9%) costs about $150 / oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Europium

For color TV tubes.

Atomic Number:	63
Atomic Symbol:	Eu
Atomic Weight:	151.96
Electron Configuration:	[Xe]$6s^2 4f^7$

History

(Europe) In 1890 Boisbaudran obtained basic fractions from samarium-gadolinium concentrates which had spark spectral lines not accounted for by samarium or gadolinium. These lines subsequently have been shown to belong to europium. The discovery of europium is generally credited to Demarcay, who separated the rare earth in reasonably pure form in 1901. The pure metal was not isolated until recent years.

Production

Europium is now prepared by mixing Eu_2O_3 with a 10%-excess of lanthanum metal and heating the mixture in a tantalum crucible under high vacuum. The element is collected as a silvery-white metallic deposit on the walls of the crucible.

Properties

As with other rare-earth metals, except for lanthanum, europium ignites in air at about 150 to 180$_o$C. Europium is about as hard as lead and is quite ductile. It is the most reactive of the rare-earth metals, quickly oxidizing in air. It resembles calcium in its reaction with water. Bastnasite and monazite are the principal ores containing europium.

Sources

Europium has been identified spectroscopically in the sun and certain stars. Seventeen isotopes are now recognized. Europium isotopes are good neutron absorbers and are being studied for use in nuclear control applications.

Uses

Europium oxide is now widely used as a phospor activator and europium-activated yttrium vanadate is in commercial use as the red phosphor in color TV tubes. Europium-doped plastic has been used as a laser material. With the development of ion-exchange techniques and special processes, the cost of the metal has been greatly reduced in recent years.

Cost

Europium is one of the rarest and most costly of the rare-earth metals. It is priced about about $7500 / kg.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Praseodymium

For ceramic coloring.

Atomic Number:	59
Atomic Symbol:	Pr
Atomic Weight:	140.9077
Electron Configuration:	$[Xe]6s^2 4f^3$

History

(Gr. prasios, green, and didymos, twin) In 1841 Mosander extracted the rare earth didymia from lanthana; in 1879, Lecoq de Boisbaudran isolated a new earth, samaria, from didymia obtained from the mineral samarskite. Six years later, in 1885, von Welsbach separated didymia into two others, praseodymia and neodymia, which gave salts of different colors. As with other rare earths, compounds of these elements in solution have distinctive sharp spectral absorption bands or lines, some of which are only a few Angstroms wide.

Sources

The element occurs along with other rare-earth elements in a variety of minerals. Monazite and bastnasite are the two principal commercial sources of the rare-earth metals. It was prepared in relatively pure form in 1931.

Production

Ion-exchange and solvent extraction techniques have led to much easier isolation of the rare earths and the cost has dropped greatly in the past few years. Praseodymium can be prepared by several methods, such as by calcium reduction of the anhydrous chloride of fluoride.

Uses

Misch metal, used in making cigarette lighters, contains about 5% praseodymium metal. The rare-earth oxides, including Pr_2O_3 are among the most refractory substances known. Along with other rare earths, it is widely used as a core material for carbon arcs used by the motion picture industry for studio lighting and projection. Salts of praseodymium are used to color glasses and enamels; when mixed with certain other materials, praseodymium produces an intense and unusually clean yellow color in glass. Didymium glass, of which praseodymium is a component,

is a colorant for welders goggles.

Properties

Praseodymium is soft, silvery, malleable, and ductile. It is somewhat more resistant to corrosion in air than europium, lanthanum, cerium, or neodymium, but it does develop a green oxide coating that spalls off when exposed to air. As with other rare-earth metals, it should be kept under a light mineral oil or sealed in plastic.

Cost

The metal (99%+ pure) is priced at about $70/oz.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Neodymium

For high strength magnets for disk drives.

Atomic Number:	60
Atomic Symbol:	Nd
Atomic Weight:	144.24
Electron Configuration:	[Xe]$6s^2 4f^4$

History

(Gr. neos, new, and didymos, twin) In 1841, Mosander, extracted from cerite a new rose-colored oxide, which he believed contained a new element. He named the element didymium, as it was an inseparable twin brother of lanthanum. In 1885 von Welsbach separated didymium into two new elemental components, neodymia and praseodymia, by repeated fractionation of ammonium didymium nitrate. While the free metal is in misch metal, long known and used as a pyrophoric alloy for light flints, the element was not isolated in relatively pure form until 1925. Neodymium is present in misch metal to the extent of about 18%. It is present in the minerals monazite and bastnasite, which are principal sources of rare-earth metals.

Production

The element may be obtained by separating neodymium salts from other rare earths by ion-exchange or solvent extraction techniques, and by reducing anhydrous halides such as NdF_3 with calcium metal. Other separation techniques are possible.

Properties

The metal has a bright silvery metallic luster, Neodymium is one of the more reactive rare-earth metals and quickly tarnishes in air, forming an oxide that spalls off and exposes metal to oxidation. The metal, therefore, should be kept under light mineral oil or sealed in a plastic material. Neodymium exists in two allotropic forms, with a transformation from a double hexagonal to a body-centered cubic structure taking place at 863^{o}C.

Isotopes

Natural neodymium is a mixture of seven stable isotopes. Fourteen other radioactive isotopes are recognized.

Uses

Didymium, of which neodymium is a component, is used for coloring glass to make welders goggles. By itself, neodymium colors glass delicate shades ranging from pure violet through wine-red and warm gray. Light transmitted through such glass shows unusually sharp absorption bands. The glass has been used in astronomical work to produce sharp bands by which spectral lines may be calibrated. Glass containing neodymium can be used as a laser material to produce coherent light. Neodymium salts are also used as a colorant for enamels.

Cost

The price of the metal is about $1/g.

Handling

Neodymium has a low-to-moderate acute toxic rating. As with other rare earths, neodymium should be handled with care.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Promethium

For nuclear batteries.

Atomic Number:	61
Atomic Symbol:	Pm
Atomic Weight:	145
Electron Configuration:	$[Xe]6s^2 4f^5$

History

(Prometheus, who, according to mythology, stole fire from heaven) In 1902 Branner predicted the existence of an element between neodymium and samarium, and this was confirmed by Moseley in 1914. In 1941, workers at Ohio State University irradiated neodymium and praseodymium with neutrons, deuterons, and alpha particles, and produced several new radioactivities, which most likely were those of element 61. Wu and Segre, and Bethe, in 1942, confirmed the formation; however, chemical proof of the production of element 61 was lacking because of the difficulty in separating the rare earths from each other at that time. In 1945, Marinsky, Glendenin, and Coryell made the first chemical identification by use of ion-exchange chromatography. Their work was doen by fission of uranium and by neutron bombardment of neodymium.

Sources

Searches for the element on earth have been fruitless, and it now appears that promethium is completely missing from the earth's crust. Promethium, however, has been identified in the spectrum of the star HR465 in Andromeda. This element is being formed recently near the star's surface, for no known isotope of promethium has a half-life longer than 17.7 years. Seventeen isotopes of promethium, with atomic masses from 134 to 155 are now known. Promethium-147, with a half-life of 2.6 years, is the most generally useful. Promethium-145 is the longest lived, and has a specific activity of 940 Ci/g.

Properties

It is a soft beta emitter; although no gamma rays are emitted, X-radiation can be generated when beta particles impinge on elements of a high atomic number, and great care must be taken in handling it. Promethium salts luminesce in the dark with a pale blue or greenish glow, due to their high radioactivity. Ion-exchange methods led to the preparation of about 10 g of premethium from atomic reactor fuel processing wastes in early 1963. Little is yet generally

known about the properties of metallic promethium. Two allotropic modifications exist.

Uses

The element has applications as a beta source for thickness gages, and it can be absorbed by a phosphor to produce light. Light produced in this manner can be used for signs or signals that require dependable operation; it can be used as a nuclear-powered battery by capturing light in photocells which convert it into electric current. Such a battery, using 147Pm, would have a useful life of about 5 years. Promethium shows promise as a portable X-ray source, and it may become useful as a heat source to provide auxiliary power for space probes and satellites. More than 30 promethium compounds have been prepared. Most are colored.

Cost

Promethium-147 is available at a cost of about 50c/Ci.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Samarium

For lasers.

Atomic Number:	62
Atomic Symbol:	Sm
Atomic Weight:	150.4
Electron Configuration:	$[Xe]6s^2 4f^6$

History

(Samarskite, a mineral) Discovered spectroscopically by its sharp absorption lines in 1879 by Lecoq de Boisbaudran in the mineral samarskite, named in honor of a Russian mine official, Col Samarski.

Sources

Samarium is found along with other members of the rare-earth elements in many minerals, including monazite and bastnasite, which are commercial sources. It occurs in monazite to the extent of 2.8%. While misch metal containing about 1% of samarium metal, has long been used, samarium has not been isolated in relatively pure form until recent years. Ion-exchange and solvent extraction techniques have recently simplified separation of the rare earths from one another; more recently, electrochemical deposition, using an electrolytic solution of lithium citrate and a mercury electrode, is said to be a simple, fast, and highly specific way to separate the rare earths. Samarium metal can be produced by reducing the oxide with lanthanum.

Properties

Samarium has a bright silver luster and is reasonably stable in air. Three crystal modifications of the metal exist, with transformations at 734 and 922_{o}C. The metal ignites in air at about 150_{o}C. The sulfide has excellent high-temperature stability and good thermoelectric efficiencies up to 1100_{o}C.

Isotopes

Twenty one isotopes of samarium exist. Natural samarium is a mixture of several isotopes, three of which are unstable with long half-lives.

Uses

Samarium, along with other rare earths, is used for carbon-arc lighting for the motion picture industry. $SmCo_5$ has been used in making a new permanent magnet material with the highest resistance to demagnetization of any known material. It is said to have an intrinsic coercive force as high as 2200 kA/m. Samarium oxide has been used in optical glass to absorb the infrared. Samarium is used to dope calcium fluoride crystal for use in optical lasers or lasers. Compounds of the metal act as sensitizers for phosphors excited in the infrared; the oxide exhibits catalytic properties in the dehydration and dehydrogenation of ethyl alcohol. It is used in infrared absorbing glass and as a neutron absorber in nuclear reactors.

Cost

The metal is priced at about $5/g.

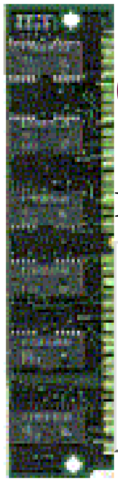
Handling

Little is known of the toxicity of samarium; therefore, it should be handled carefully.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated:12/19/97, CST Information Services Team

Periodic Table



Gadolinium

For computer memory.

Atomic Number:	64
Atomic Symbol:	Gd
Atomic Weight:	157.25
Electron Configuration:	$[Xe]6s^2 4f^7 5d^1$

History

From gadolinite, a mineral named for Gadolin, a Finnish chemist. The rare earth metal is obtained from the mineral *gadolinite*. Gadolinia, the oxide of gadolinium, was separated by Marignac in 1880 and Lecoq de Boisbaudran independently isolated it from Mosander's *yttria* in 1886.

Sources

Gadolinium is found in several other minerals, including monazite and bastnasite, both of which are commercially important. With the development of ion-exchange and solvent extraction techniques, the availability and prices of gadolinium and the other rare-earth metals have greatly improved. The metal can be prepared by the reduction of the anhydrous fluoride with metallic calcium.

Isotopes

Natural gadolinium is a mixture of seven isotopes, but 17 isotopes of gadolinium are now recognized. Although two of these, 155Gd and 157Gd, have excellent capture characteristics, they are only present naturally in low concentrations. As a result, gadolinium has a very fast burnout rate and has limited use as a nuclear control rod material.

Properties

As with other related rare-earth metals, gadolinium is silvery white, has a metallic luster, and is malleable and ductile. At room temperature, gadolinium crystallizes in the hexagonal, close-packed alpha form. Upon heating to 1235$_o$C, alpha gadolinium transforms into the beta form, which has a body-centered cubic structure.

The metal is relatively stable in dry air, but tarnishes in moist air and forms a losely adhering

oxide film which spalls off and exposes more surface to oxidation. The metal reacts slowly with water and is soluble in dilute acid.

Gadolinium has the highest thermal neutron capture cross-section of any known element (49,000 barns).

Uses

Gadolinium yttrium garnets are used in microwave applications and gadolinium compounds are used as phosphors in color television sets.

The metal has unusual superconductive properties. As little as 1 percent gadolinium improves the workability and resistance of iron, chromium, and related alloys to high temperatures and oxidation.

Gadolinium ethyl sulfate has extremely low noise characteristics and may find use in duplicating the performance of amplifiers, such as the maser.

The metal is ferromagnetic. Gadolinium is unique for its high magnetic movement and for its special Curie temperature (above which ferromagnetism vanishes) lying just at room temperature. This suggests applications as a magnetic component that can sense hot and cold.

Costs

Previous to 1993, the price of the metal was $485/kg.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Terbium

For fluorescent lamps.

Atomic Number:	65
Atomic Symbol:	Tb
Atomic Weight:	158.9254
Electron Configuration:	[Xe]$6s^2 4f^9$

History

(Ytterby, a village in Sweden) Discovered by Mosander in 1843. Terbium is a member of the lanthanide or "rare earth" group of elements. It is found in cerite, gadolinite, and other minerals along with other rare earths. It is recovered commercially from monazite in which it is present to the extent of 0.03%, from xenotime, and from euxenite, a complex oxide containing 1% or more of terbia.

Production

Terbium has been isolated only in recent years with the development of ion-exchange techniques for separating the rare-earth elements. As with other rare earths, it can be produced by reducing the anhydrous chloride or fluoride with calcium metal in a tantalum crucible. Calcium and tantalum impurities can be removed by vacuum remelting. Other methods of isolation are possible.

Properties

Terbium is reasonably stable in air. It is a silver-gray metal, and is malleable, ductile, and soft enough to be cut with a knife. Two crystal modifications exist, with a transformation temperature of 1289$_{o}$C. Twenty one isotopes with atomic masses ranging from 145 to 165 are recognized. The oxide is a chocolate or dark maroon color.

Uses

Sodium terbium borate is used in solid-state devices. The oxide has potential application as an activator for green phosphors used in color TV tubes. It can be used with ZrO_2 as a crystal stabilizer of fuel cells which operate at elevated temperature. Few other uses have been found.

Cost

The element is priced at about $30 / g (99.9%).

Handling

Little is known of the toxicity of terbium. It should be handled with care as with other lanthanide elements.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Dysprosium

For color TV tubes.

Atomic Number:	66
Atomic Symbol:	Dy
Atomic Weight:	162.50
Electron Configuration:	$[Xe]6s^2 4f^{10}$

History

(Gr. dysprositos, hard to get at) Dysprosium was discovered in 1886 by Lecoq de Boisbaudran, but not isolated. Neither the oxide nor the metal was available in relatively pure form until the development of ion-exchange separation and metallographic reduction techniques by Spedding and associates about 1950. Dysprosium occurs along with other so-called rare-earth or lanthanide elements in a variety of minerals such as xenotime, fergusonite, gadolinite, euxenite, polycrase, and blomstrandine. The most important sources, however, are from monaziate and bastnasite. Dysprosium can be prepared by reduction of the trifluoride with calcium.

Properties

The element has a metallic, bright silver luster. It is relatively stable in air at room temperature, and is readily attacked and dissolved, with the evolution of hydrogen, but dilute and concentrated mineral acids. The metal is soft enough to be cut with a knife and can be machined without sparking if overheating is avoided. Small amounts of impurities can greatly affect its physical properties.

Uses

While dysprosium has not yet found many applications, its thermal neutron absorption cross-section and high melting point suggest metallurgical uses in nuclear control applications and for alloying with special stainless steels. A dysprosium oxide-nickel cermet has found use in cooling nuclear reactor rods. This cermet absorbs neutrons readily without swelling or contracting under prolonged neutron bombardment. In combination with vanadium and other rare earths, dysprosium has been used in making laser materials. Dysprosium-cadmium chalcogenides, as sources of infrared radiation, have been used for studying chemical reactions.

Cost

The cost of dysprosium metal has dropped in recent years since the development of ion-exchange and solvent extraction techniques, and the discovery of large ore bodies. The metal costs about $300 / kg in purities of 99+%.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Holmium

For eye-safe lasers.

Atomic Number:	67
Atomic Symbol:	Ho
Atomic Weight:	164.9304
Electron Configuration:	$[Xe]6s^2 4f^{11}$

History

(L. Holmia, for Stockholm). The special absorption bands of holmium were noticed in 1878 by the Swiss chemists Delafontaine and Soret, who announced the existence of an "Element X." Cleve, of Sweden, later independently discovered the element while working on erbia earth. The element is named after cleve's native city. Holmia, the yellow oxide, was prepared by Homberg in 1911. Holmium occurs in gadolinite, monazite, and in other rare-earth minerals. It is commercially obtained from monazite, occurring in that mineral to the extent of about 0.05%. It has been isolated by the reduction of its anhydrous chloride or fluoride with calcium metal.

Properties

Pure holmium has a metallic to bright silver luster. It is relatively soft and malleable, and is stable in dry air at room temperature, but rapidly oxidizes in moist air and at elevated temperatures. The metal has unusual magnetic properties. Few uses have yet been found for the element. The element, as with other rare earths, seems to have a low acute toxic rating.

Cost

The price of 99+% holmium metal is about $10/g.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Erbium

For coating for sunglasses.

Atomic Number:	68
Atomic Symbol:	Er
Atomic Weight:	167.26
Electron Configuration:	$[Xe]6s^{2}4f^{12}$

History

(Ytterby, a town in Sweden) Erbium, one of the so-called rare-earth elements on the lanthanide series, is found in the minerals mentioned under dysprosium. In 1842 Mosander separated "yttria" found in the mineral gadolinite, into three fractions which he called yttria, erbia, and terbia. The names erbia and terbia became confused in this early period. After 1860, Mosander's terbia was known as erbia, and after 1877, the earlier known erbia became terbia. The erbia of this period was later shown to consist of five oxides, now known as erbia, scandia, holmia, thulia and ytterbia. By 1905 Urbain and James independently succeeded in isolating fairly pure Er_2O_3. Klemm and Bommer first produced reasonably pure erbium metal in 1934 by reducing the anhydrous chloride with potassium vapor. The pure metal is soft and malleable and has a bright, silvery, metallic luster. As with other rare-earth metals, its properties depend to a certain extent on the impurities present. The metal is fairly stable in air and does not oxidize as rapidly as some of the other rare-earth metals. Naturally occurring erbium is a mixture of six isotopes, all of which are stable. Nine radioactive isotopes of erbium are also recognized. Recent production techniques, using ion-exchange reactions, have resulted in much lower prices of the rare-earth metals and their compounds in recent years. The cost of 99+% erbium metal is about $650/kg. Erbium is finding nuclear and metallurgical uses. Added to vanadium, for example, erbium lowers the hardness and improves workability. Most of the rare-earth oxides have sharp absorption bands in the visible, ultraviolet, and near infrared. This property, associated with the electronic structure, gives beautiful pastel colors to many of the rare-earth salts. Erbium oxide gives a pink color and has been used as a colorant in glasses and porcelain enamel glazes.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Thulium

For lasers.

Atomic Number:	69
Atomic Symbol:	Tm
Atomic Weight:	168.9342
Electron Configuration:	$[Xe]6s^2 4f^{13}$

History

(Thule, the earliest name for Scandinavia) Discovered in 1879 by Cleve. Thulium occurs in small quantities along with other rare earths in a number of minerals. It is obtained commercially from monazite, which contains about 0.007% of the element. Thulium is the least abundant of the rare earth elements, but with new sources recently discovered, it is now considered to be about as rare as silver, gold, or cadmium. Ion-exchange and solvent extraction techniques have recently permitted much easier separation of the rare earths, with much lower costs. Only a few years ago, thulium metal was not obtainable at any cost; in 1985 the oxide sold for $3400/kg. Thulium metal costs $50/g. Thulium can be isolated by reduction of the oxide with lanthanum metal or by calcium reduction of a closed container. The element is silver-gray, soft, malleable, and ductile, and can be cut with a knife. Twenty five isotopes are known, with atomic masses ranging from 152 to 176. Natural thulium, which is 100% 169Tm, is stable. Because of the relatively high price of the metal, thulium has not yet found many practical applications. 169Tm bombarded in a nuclear reactor can be used as a radiation source in portable X-ray equipment. 171Tm is potentially useful as an energy source. Natural thulium also has possible use in ferrites (ceramic magnetic materials) used in microwave equipment. As with other lanthanides, thulium has a low-to-moderate acute toxic rating. It should be handled with care.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Ytterbium

For dentures.

Atomic Number:	70
Atomic Symbol:	Yb
Atomic Weight:	173.04
Electron Configuration:	[Xe]$6s^2 4f^{14}$

History

(Ytterby, village in Sweden) Marignac in 1878 discovered a new component, which he called ytterbia, in the earth then known as erbia. In 1907, Urbain separated ytterbia into two components, which he called neoytterbia and lutecia. The elements in these earths are now known as ytterbium and lutetium, respectively. These elements are identical with aldebaranium and cassiopeium, discovered independently and at about the same time by von Welsbach.

Sources

Ytterbium occurs along with other rare earths in a number of rare minerals. It is commercially recovered principally from monazite sand, which contains about 0.03%. Ion-exchange and solvent extraction techniques developed in recent years have greatly simplified the separation of the rare earths from one another.

Production

The element was first prepared by Klemm and bonner in 1937 by reducing ytterbium trichloride with potassium. Their metal was mixed, however, with KCl. Daane, Dennison, and Spedding prepared a much purer from in 1953 from which the chemical and physical properties of the element could be determined.

Properties

Ytterbium has a bright silvery luster, is soft, malleable, and quite ductile. While the element is fairly stable, it should be kept in closed containers to protect it from air and moisture. Ytterbium is readily attacked and dissolved by dilute and concentrated mineral acids and reacts slowly with water. Ytterbium has three allotropic forms with transformation points at -13_oC and 795_oC. The beta form is a room-temperature, face-centered, cubic modification, while the

high-temperature gamma form is a body-centered cubic form. Another body-centered cubic phase has recently been found to be stable at high pressures at room temperatures. The beta form ordinarily has metallic-type conductivity, but becomes a semiconductor when the pressure is increased about 16,000 atm. The electrical resistance increases tenfold as the pressure is increased to 39,000 atm and drops to about 10% of its standard temperature-pressure resistivity at a pressure of 40,000 atm. Natural ytterbium is a mixture of seven stable isotopes. Seven other unstable isotopes are known.

Uses

Ytterbium metal has possible use in improving the grain refinement, strength, and other mechanical properties of stainless steel. One isotope is reported to have been used as a radiation source substitute for a portable X-ray machine where electricity is unavailable. Few other uses have been found.

Cost

Ytterbium metal is commercially available with a purity of about 99+% for about $875/kg.

Handling

Ytterbium has a low acute toxic rating.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Lutetium

For dentures.

Atomic Number:	71
Atomic Symbol:	Lu
Atomic Weight:	174.97
Electron Configuration:	$[Xe]6s^2 4f^{14} 5d^1$

History

(Lutetia, ancient name for Paris, sometimes called cassiopeium by the Germans) In 1907, Urbain described a process by which Marignac's ytterbium (1879) could be separated into the two elements, ytterbium (neoytterbium) and lutetium. These elements were identical with "aldebaranium" and "cassiopeium," independently discovered at this time. The spelling of the element was changed from lutecium to lutetium in 1949. Lutetium occurs in very small amounts in nearly all minerals containing yttrium, and is present in monazite to the extent of about 0.003%, which is a commercial source. The pure metal has been isolated only in recent years and is one of the most difficult to prepare. It can be prepared by the reduction of anhydrous $LuCl_3$ or LuF_3 by an alkali or alkaline earth metal. The metal is silvery white and relatively stable in air. While new techniques, including ion-exchange reactions, have been developed to separate the various rare-earth elements, lutetium is still the most costly of all rare earths. It is priced at about $75 / g. 176Lu occurs naturally (2.6%) with 175Lu (97.4%). It is radioactive with a half-life of about 3×10^{10} years. Stable lutetium nuclides, which emit pure beta radiation after thermal neutron activation, can be used as catalysts in cracking, alkylation, hydrogenation, and polymerization. Virtually no other commercial uses have been found yet for lutetium. While lutetium, like other rare-earth metals, is thought to have a low toxicity rating, it should be handled with care until more information is available.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Protactinium

Atomic Number:	91
Atomic Symbol:	Pa
Atomic Weight:	231.0359
Electron Configuration:	$[Rn]7s^2 5f^2 6d^1$

History

(Gr. protos, first) The first isotope of element 91 to be discovered was 234Pa, also known as UX2, a short-lived member of the naturally occurring 238U decay series. It was identified by K. Fajans and O.H. Gohring in 1913 and the named the new element brevium. When the longer-lived isotope 231-Pa was identified by Hahn and Meitner in 1918, the name protoactinium was adopted as being more consistent with the characteristics of the most abundant isotope. Sody, Cranson, and Fleck were also active in this work. The name protoactinium was shortened to protactinium in 1949. In 1927, Grosse prepared 2 mg of a white powder, which was shown to be Pa2O5. Later, in 1934, from 0.1 g of pure Pa2O5 he isolated the element by two methods, one of which was by converting the oxide to an iodide and "cracking" it in a high vacuum by an electrically heated filament by the reaction: $2PaI_5$ --> 2Pa + $5I_2$. Protactinium has a bright metallic luster which it retains for some time in air. The element occurs in pitchblende to the extent of about 1 part 231Pa to 10 million of ore. Ores from Zaire have about 3 ppm. Protactinium has 20 isotopes, the most common of which is 231Pa with a half-life of 32,700 years. A number of protactinium compounds are known, some of which are colored. The element is superconductive below 1.4K. The element is a dangerous material and requires precautions similar to those used when handling plutonium. In 1959 and 1961, it was announced that the Great Britain Atomic Energy Authority extracted by a 12-stage process 125 g of 99.9% protactinium, the world's only stock of the metal for many years to come. The extraction was made from 60 tons of waste material at a cost of about $500,000. Protactinium is one of the rarest and most expensive naturally occurring elements. O.R.N.L. supplies promethium-231 at a cost of about $280/g. The elements is an alpha emitter (5.0 MeV) and is a radiological hazard similar to polonium.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Uranium

For Nuclear Fission

Atomic Number:	92
Atomic Symbol:	U
Atomic Weight:	238.029
Electron Configuration:	$[Rn]7s^2 5f^3 6d^1$

History

(Planet Uranus) Yellow-colored glass, containing more than 1% uranium oxide and dating back to 79 A.D., has been found near Naples, Italy. Klaproth recognized an unknown element in pitchblende and attempted to isolate the metal in 1789.

The metal apparently was first isolated in 1841 by Peligot, who reduced the anhydrous chloride with potassium.

Sources

Uranium, not as rare as once thought, is now considered to be more plentiful than mercury, antimony, silver, or cadmium, and is about as abundant as molybdenum or arsenic. It occurs in numerous minerals such as pitchblende, uraninite, carnotite, autunite, uranophane, and tobernite. It is also found in phosphate rock, lignite, monazite sands, and can be recovered commercially from these sources.

The United States Department of Energy purchases uranium in the form of acceptable U_3O_8 concentrates. This incentive program has greatly increased the known uranium reserves.

Uranium can be prepared by reducing uranium halides with alkali or alkaline earth metals or by reducing uranium oxides by calcium, aluminum, or carbon at high temperatures. The metal can also be produced by electrolysis of KUF_5 or UF_4, dissolved in a molten mixture of $CaCl_2$ and NaCl. High-purity uranium can be prepared by the thermal decomposition of uranium halides on a hot filament.

Properties

Uranium exhibits three crystallographic modifications as follows: alpha --(688C)--> beta --(776C)--> gamma. Uranium is a heavy, silvery-white metal which is pyrophoric when finely divided.

It is a little softer than steel, and is attacked by cold water in a finely divided state. It is malleable, ductile, and slightly paramagnetic.

In air, the metal becomes coated with a layer of oxide. Acids dissolve the metal, but it is unaffected by alkalis.

Isotopes

Uranium has sixteen isotopes, all of which are radioactive. Naturally occurring uranium nominally contains 99.28305 by weight ^{238}U, 0.7110% ^{235}U, and 0.0054% ^{234}U. Studies show that the percentage weight of ^{235}U in natural uranium varies by as much as 0.1%, depending on the source. The US DOE has adopted the value of 0.711 as being their *official* percentage of ^{235}U in natural uranium. Natural uranium is sufficiently radioactive to expose a photographic plate in an hour or so.

Much of the internal heat of the earth is thought to be attributable to the presence of uranium and thorium.

Uranuim-238 with a half-life of 4.51 x 10^9 years, has been used to estimate the age of igneous rocks. The origin of uranium, the highest member of the naturally occurring elements - except perhaps for traces of neptunium or plutonium, is not clearly understood. However it may be presumed that uranium is a decay product of elements with higher atomic weight, which may have once been present on earth or elsewhere in the universe. These original elements may have been formed as a result of a primordial *creation*, known as *the big bang*, in a supernova, or in some other stellar processes.

Uses

Uranium is of great importance as a nuclear fuel. Uranium-238 can be converted into fissionable plutonium by the following reactions: ^{238}U(n, gamma) --> ^{239}U --(beta)--> ^{239}Np --(beta)--> ^{239}Pu. This nuclear conversion can be brought about in breeder reactors where it is possible to produce more new fissionable material than the fissionable material used in maintaining the chain reaction.

Uranium-235 is of even greater importance because it is the key to utilizing uranium. ^{235}U, while occuring in natural uranium to the extent of only 0.71%, is so fissionable with slow neutrons that a self-sustaining fission chain reaction can be made in a reactor constructed from natural uranium and a suitable moderator, such as heavy water or graphite, alone.

Uranium-235 can be concentrated by gaseous diffusion and other physical processes, if desired, and used directly as a nuclear fuel, instead of natural uranium, or used as an explosive.

Natural uranium, slightly enriched with ^{235}U by a small percentage, is used to fuel nuclear power reactors to generate electricity. Natural thorium can be irradiated with neutrons as follows to produce the important isotope ^{233}U: ^{232}Th(n, gamma)--> ^{233}Th --(beta)--> ^{233}Pa --(beta)--> ^{233}U. While thorium itself is not fissionable, ^{233}U is, and in this way may be used as a

nuclear fuel. One pound of completely fissioned uranium has the fuel value of over 1500 tons of coal.

The uses of nuclear fuels to generate electrical power, to make isotopes for peaceful purposes, and to make explosives are well known. The estimated world-wide capacity of the 429 nuclear power reactors in operation in January 1990 amounted to about 311,000 megawatts.

Uranium in the U.S.A. is controlled by the U.S. Nuclear Regulatory Commission. New uses are being found for depleted uranium, ie., uranium with the percentage of ^{235}U lowered to about 0.2%.

Uranium is used in inertial guidance devices, in gyro compasses, as counterweights for aircraft control surfaces, as ballast for missile reentry vehicles, and as a shielding material. Uranium metal is used for X-ray targets for production of high-energy X-rays; the nitrate has been used as a photographic toner, and the acetate is used in analytical chemistry.

Crystals of uranium nitrate are triboluminescent. Uranium salts have also been used for producing yellow "vaseline" glass and glazes. Uranium and its compounds are highly toxic, both from a chemical and radiological standpoint.

Handling

Finely divided uranium metal, being pyrophoric, presents a fire hazard.

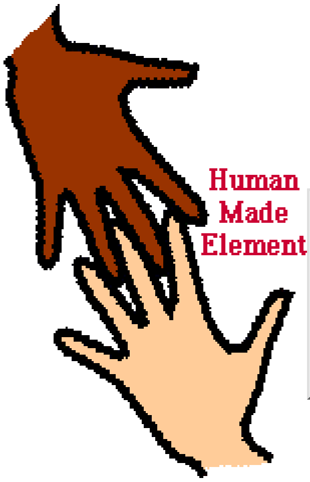
Working with uranium requires the knowledge of the maximum allowable concentrations that may be inhaled or ingested.

Recently, the natural presence of uranium in many soils has become of concern to homeowners because of the generation of radon and its daughters.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Neptunium

Atomic Number:	93
Atomic Symbol:	Np
Atomic Weight:	237.0482
Electron Configuration:	$[Rn]7s^2 5f^4 6d^1$

History

(Planet Neptune) Neptunium was the first synthetic transuranium element of the actinide series discovered; the isotope ^{239}Np was produced by McMillan and Abelson in 1940 at Berkeley, California, as the result of bombarding uranium with cyclotron-produced neutrons. The isotope ^{237}Np (half-life of 2.14×10^6 years) is currently obtained in gram quantities as a by-product from nuclear reactors in the production of plutonium. Trace quantities of the element are actually found in nature due to transmutation reactions in uranium ores produced by the neutrons which are present. Neptunium is prepared by the reduction of NpF_3 with barium or lithium vapor at about 1200C. Neptunium metal has a silvery appearance, is chemically reactive, and exists in at least three structural modifications: alpha-neptunium, orthorhombic, density 20.25 g/cm^3, beta-neptunium (above 280C), tetragonal, density (313C) 19.36 g/cm^3, gamma-neptunium (above 577C), cubic, density (600C) 18.0 g/cm^3. Neptunium has four ionic oxidation states in solution: Np^{+3} (pale purple), analogous to the rare earth ion Pm^{+3}, Np^{+4} (yellow green); NpO^{+} (green blue): and NpO^{++} (pale pink). These latter oxygenated species are in contrast to the rare earths which exhibit only simple ions of the (II), (III), and (IV) oxidation states in aqueous solution. The element forms tri- and tetrahalides such as NpF_3, NpF_4, $NpCl_4$, $NpBr_3$, NpI_3, and oxides of the various compositions such as are found in the uranium-oxygen system, including Np_3O_8 and NpO_2. Fifteen isotopes of neptunium are now recognized. The O.R.N.L. has ^{237}Np available for sale to its licensees and for export. This isotope can be used as a component in neutron detection instruments. It is offered at a price of $280/g.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Plutonium

For nuclear batteries.

Atomic Number:	94
Atomic Symbol:	Pu
Atomic Weight:	244
Electron Configuration:	[Rn]$7s^2 5f^6$

History

(Planet pluto) Plutonium was the second transuranium element of the actinide series to be discovered. The isotope ^{238}Pu was produced in 1940 by Seaborg, McMillan, Kennedy, and Wahl by deuteron bombardment of uranium in the 60-inch cyclotron at Berkeley, California. Plutonium also exists in trace quantities in naturally occurring uranium ores. It is formed in much the same manner as neptunium, by irradiation of natural uranium with the neutrons which are present.

Isotopes

By far of greatest importance is the isotope Pu^{239}, with a half-life of 24,100 years, produced in extensive quantities in nuclear reactors from natural uranium: ^{238}U(n, gamma) --> ^{239}U--(beta) --> ^{239}Np--(beta) --> ^{239}Pu. Fifteen isotopes of plutonium are known.

Plutonium also exhibits four ionic valence states in aqueous solutions: Pu^{+3} (blue lavender), Pu^{+4} (yellow brown), PuO^{+} (pink?), and PuO^{+2}(pink-orange). The ion PuO^{+} is unstable in aqueous solutions, disproportionating into Pu^{+4} and PuO^{+2}. The Pu^{+4} thus formed, however, oxidizes the PuO^{+} into PuO^{+2}, itself being reduced to Pu^{+3}, giving finally Pu^{+3} and PuO^{+2}. Plutonium forms binary compounds with oxygen: PuO, PuO_2, and intermediate oxides of variable composition; with the halides: PuF_3, PuF_4, $PuCl_3$, $PuBr_3$, PuI_3; with carbon, nitrogen, and silicon: PuC, PuN, $PuSi_2$. Oxyhalides are also well known: PuOCl, PuOBr, PuOI.

Uses

Plutonium has assumed the position of dominant importance among the trasuranium elements because of its successful use as an explosive ingredient in nuclear weapons and the place which it holds as a key material in the development of industrial use of nuclear power. One kilogram is equivalent to about 22 million kilowatt hours of heat energy. The complete detonation of a kilogram of plutonium produces an explosion equal to about 20,000 tons of chemical explosive.

Its importance depends on the nuclear property of being readily fissionable with neutrons and its availability in quantity. The world's nuclear-power reactors are now producing about 20,000 kg of plutonium / yr. By 1982 it was estimated that about 300,000 kg had accumulated. The various nuclear applications of plutonium are well known. 238Pu has been used in the Apollo lunar missions to power seismic and other equipment on the lunar surface. As with neptunium and uranium, plutonium metal can be prepared by reduction of the trifluoride with alkaline-earth metals.

Properties

The metal has a silvery appearance and takes on a yellow tarnish when slightly oxidized. It is chemically reactive. A relatively large piece of plutonium is warm to the touch because of the energy given off in alpha decay. Larger pieces will produce enough heat to boil water. The metal readily dissolves in concentrated hydrochloric acid, hydroiodic acid, or perchloric acid. The metal exhibits six allotropic modifications having various crystalline structures. The densities of these vary from 16.00 to 19.86 g / cm^3.

Hazards

Because of the high rate of emission of alpha particles and the element being specifically absorbed on bone the surface and collected in the liver, plutonium, as well as all of the other transuranium elements except neptunium, are radiological poisons and must be handled with very special equipment and precautions. Plutonium is a very dangerous radiological hazard. Precautions must also be taken to prevent the unintentional formulation of a critical mass. Plutonium in liquid solution is more likely to become critical than solid plutonium. The shape of the mass must also be considered where criticality is concerned.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Americium

For crystal research.

Atomic Number:	95
Atomic Symbol:	Am
Atomic Weight:	243
Electron Configuration:	$[Rn]7s^{2}5f^{7}$

History

(the Americas) Americium was the fourth transuranic element to be discovered; the isotope 241Am was identified by Seaborg, James, Morgan, and Ghiorso late in 1944 at the wartime Metallurgical Laboratory of the University of Chicago as the result of successive neutron capture reactions by plutonium isotopes in a nuclear reactor. The luster of freshly prepared americium metal is white and more silvery than plutonium or neptunium prepared in the same manner. It appears to be more malleable than uranium or neptunium and tarnishes slowly in dry air at room temperature. Americium must be handled with great care to avoid personal contamination. The alpha activity from 241Am is about three times that of radium. When gram quantities of 241Am are handled, the intense gamma activity makes exposure a serious problem. 241Am has been used as a portable source for gamma radiography. It has also been used as a radioactive glass thickness gauge for the flat glass industry and as a source of ionization for smoke detectors.

Isotope

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table

Curium

Atomic Number:	96
Atomic Symbol:	Cm
Atomic Weight:	247
Electron Configuration:	$[Rn]7s^2 5f^7 6d^1$

History

(Pierre and Marie Curie) Although curium follows americium in the periodic system, it was actually known before americium and was the third transuranium element to be discovered. It was identified by Seaborg, James, and Ghiorso in 1944 at the wartime Metallurgical Laboratory in Chicago as a result of helium-ion bombardment of 239Pu in the Berkeley, California, 60-inch cyclotron. Visible amounts (30Mg) of 242Cm, in the form of the hydroxide, were first isolated by Werner and Perlman of the University of California in 1947. In 1950, Crane, Wallmann, and Cunningham found that the magnetic susceptibility of microgram samples of CmF3 was of the same magnitude as that of GdF3. This provided direct experimental evidence for assigning an electronic configuration to Cm^{+3}. In 1951, the same workers prepared curium in its elemental form for the first time. Fourteen isotopes of curium are now known. The most stable, 247Cm, with a half-life of 16 million years, is so short compared to the earth's age that any primordial curium must have disappeared long ago from the natural scene. Minute amounts of curium probably exist in natural deposits of uranium, as a result of a sequence of neutron captures and beta decays sustained by the very low flux of neutrons naturally present in uranium ores. The presence of natural curium, however, has never been detected. 242Cm and 244Cm are available in multigram quantities. 248Cm has been produced only in milligram amounts. Curium is similar in some regards to gadolinium, its rare earth homolog, but it has a more complex crystal structure. Curium is silver in color, is chemically reactive, and is more electropositive than aluminum. Most compounds of trivalent curium are faintly yellow in color. 242 Cm generates about three watts of thermal energy per gram. This compares to one-half watt per gram of 238Pu. This suggests use for curium as a power source. 244Cm is now offered for sale at $100/mg. Curium absorbed into the body accumulates in the bones, and is therefore very toxic as its radiation destroys the red-cell forming mechanism. The maximum permissible total body burden of 244Cm (soluble) in a human being is 0.3 microcurie.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Berkelium

Atomic Number:	97
Atomic Symbol:	Bk
Atomic Weight:	247
Electron Configuration:	$[Rn]7s^2 5f^9$

History

(Berkeley, home of the University of California) Berkelium, the eighth member of the actinide transition series, was discovered in December 1949 by Thompson, Ghiorso, and Seaborg, and was the fifth transuranium element synthesized. It was produced by cyclotron bombardment of milligram amounts of 241Am with helium ions at Berkeley, California. The first isotope produced had a mass of 243 and decayed with a half-life of 4.5 hours. Ten isotopes are now known and have been synthesized. The evidence of 249Bk with a half-life of 314 days, makes it feasible to isolate berkelium in weighable amounts so that its properties can be investigated with macroscopic quantities. One of the first visible amounts of a pure berkelium compound, berkelium chloride, was produced in 1962. It weighed 1 billionth of a gram. Berkelium probably has not yet been prepared in elemental form, but is expected to be a silvery metal, easily soluble in dilute mineral acids, and readily oxidized by air or oxygen at elevated temperatures to form the oxide. X-ray diffraction methods have been used to identify various compounds. As with other actinide elements, berkelium tends to accumulate in the skeletal system. Because of its rarity, berkelium presently has NO COMMERCIAL OR TECHNOLOGICAL USE.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Californium

For rocket fuel

Atomic Number:	98
Atomic Symbol:	Cf
Atomic Weight:	251
Electron Configuration:	[Rn]$7s^{2}5f^{10}$

History

(State and University of California) Californium, the sixth transuranium element to be discovered, was produced by Thompson, Street, Ghioirso, and Seaborg in 1950 by bombarding microgram quantities of 242Cm with 35 MeV helium ions in the Berkeley 60-inch cyclotron. Californium (III) is the only ion stable in aqueous solutions, all attempts to reduce or oxidize californium (III) having failed. The isotope 249Cf results from the beta decay of 249Bk while the heavier isotopes are produced by intense neutron irradiation by the reactions. The existence of the isotopes 249Cf, 250Cf, 251Cf, and 252Cf makes it feasible to isolate californium in weighable amounts so that its properties can be investigated with macroscopic quantities. Californium-252 is a very strong neutron emitter. One microgram releases 170 million neutrons per minute, which presents biological hazards. Proper safeguards should be used in handling californium. Reduction of californium to its metallic state has not yet been accomplished. Because californium is a very efficient source of neutrons, many new uses are expected for it. It has already found use in neutron moisture gauges and in well-logging (the determination of water and oil-bearing layers). It is also being used as a portable neutron source for discovery of metals such as gold or silver by on-the-spot activation analysis. 252-Cf is now being offered for sale by the O.R.N.L. at a cost of $10/mg. As of May, 1975, more than 63 mg have been produced and sold. It has been suggested that californium may be produced in certain stellar explosions, called supernovae, for the radioactive decay of 254Cf (55-day half-life) agrees with the characteristics of the light curves of such explosions observed through telescopes. This suggestion, however, is questioned.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Einsteinium

Atomic Number:	99
Atomic Symbol:	Es
Atomic Weight:	252
Electron Configuration:	$[Rn]7s^{2}5f^{11}$

History

(Albert Einstein) Einsteinium, the seventh transuranic element of the actinide series to be discovered, was identified by Ghiorso and co-workers at Berkeley in December 1952 in debris from the first large thermonuclear explosion, which took place in the Pacific in November, 1952. The 20-day 253Es isotope was produced.

In 1961, enough einsteinium was produced to separate a macroscopic amount of 253Es. This sample weighted about 0.01Mg and was measured using a special magnetic-type balance. 253Es so produced was used to produce mendelevium (Element 101).

About 3Mg of einsteinium has been produced at Oak Ridge National Laboratories by

1. irridiating kilogram quantities of 239Pu in a reactor for several years to produce 242Pu,
2. fabricating the 242Pu into pellets of plutonium oxide and aluminum powder,
3. loading the pellets into target rods for an initial 1-year irradiation at the Savannah River Plant, and
4. irradiating the targets for another 4 months in a HFIR (High Flux Isotopic Reactor).

The targets were then removed for chemical separation of the einsteinium from californium.

Isotopes

Fourteen isotopes of einsteinium are now recognized. 254Es has the longest half-life (275 days).

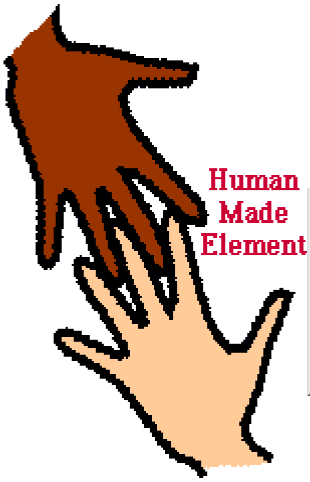
Properties

Tracer studies using 253Es show that einsteinium has chemical properties typical of a heavy trivalent, actinide element.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, *CST Information Services Team*

Periodic Table



Fermium

Atomic Number:	100
Atomic Symbol:	Fm
Atomic Weight:	257
Electron Configuration:	-30-8-2

History

(Enrico Fermi) Fermium, the eighth discovered transuranium element of the actinide series, was identified by Ghiorso and co-workers in 1952 in the debris from a thermonuclear explosion in the pacific during work involving the University of California Radiation Laboratory, Argonne National Laboratory, and Los Alamos Scientific Laboratory.

The isotope produced was the 20-hour 255Fm. During 1953 and early 1954, while discovery of elements 99 and 100 was withheld from publication for security reasons, a group from the Nobel Institute of Physics in Stockholm bombarded 238U with 16O ions, and isolated a 30-min alpha-emitter, which they ascribed to 250-100, without claiming discovery of the element. This isotope has since been identified positively, and the 30-min half-life confirmed.

Properties

The chemical properties of fermium have been studied solely with tracer amounts. In normal aqueous media, only the (III) oxidation state appears to exist.

Isotopes

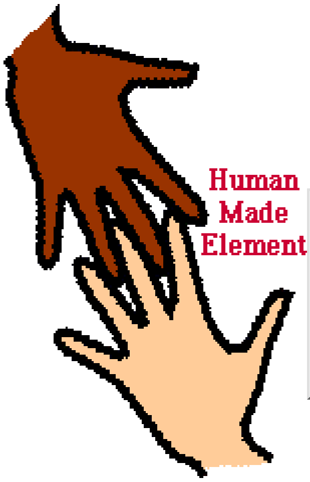
254Fm and heavier isotopes can be produced by intense neutron irradiation of lower elements, such as plutonium, using a process of successive neutron capture interspersed with beta decays until these mass numbers and atomic numbers are reached.

Sixteen isotopes of fermium are known to exist. 257Fm, with a half-life of about 100.5 days, is the longest lived. 250Fm, with a half-life of 30 minutes, has been shown to be a decay product of element 254-102. Chemical identification of 250Fm confirmed the production of element 102 (nobelium).

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table



Mendelevium

Atomic Number:	101
Atomic Symbol:	Md
Atomic Weight:	258
Electron Configuration:	$[Rn]7s^2 5f^{13}$

History

(Dmitri Mendeleev) Mendelevium, the ninth transuranium element of the actinide series discovered, was first identified by Ghiorso, Harvey, Choppin, Thompson, and Seaborg in early in 1955 during the bombardment of the isotope 253Es with helium ions in the Berkeley 60-inch cyclotron. The isotope produced was 256Md, which has a half-life of 76 min. This first identification was notable in that 256Md was synthesized on a one-atom-at-a-time basis.

Isotopes

Fourteen isotopes are now recognized. 258Md has a half-life of 2 months. This isotope has been produced by the bombardment of an isotope of einsteinium with ions of helium. Eventually enough 258Md should be made to determine its physical properties.

Uses

256Md has been used to elucidate some of the chemical properties of mendelevium in aqueous solution.

Properties

Experiments seem to show that the element possesses a moderately stable dipositive (II) oxidation state in addition to the tripositive (III) oxidation state, which is characteristic of the actinide elements.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97, CST Information Services Team

Periodic Table

Lawrencium

Atomic Number:	103
Atomic Symbol:	Lr
Atomic Weight:	262
Electron Configuration:	[Rn]$7s^2 5f^{14} 6d^1$

History

(Ernest O. Lawrence, inventor of the cyclotron) This member of the 5f transition elements (actinide series) was discovered in March 1961 by A. Ghiorso, T. Sikkeland, A.E. Larsh, and R.M. Latimer. A 3-Mg californium target, consisting of a mixture of isotopes of mass number 249, 250, 251, and 252, was bombarded with either 10B or 11B. The electrically charged transmutation nuclei recoiled with an atmosphere of helium and were collected on a thin copper conveyor tape which was then moved to place collected atoms in front of a series of solid-state detectors. The isotope of element 103 produced in this way decayed by emitting an 8.6 MeV alpha particle with a half-life of 8 s.

In 1967, Flerov and associates at the Dubna Laboratory reported their inability to detect an alpha emitter with a half-life of 8 s which was assigned by the Berkeley group to 257-103. This assignment has been changed to 258Lr or 259Lr.

In 1965, the Dubna workers found a longer-lived lawrencium isotope, 256Lr, with a half-life of 35 s. In 1968, Thiorso and associates at Berkeley used a few atoms of this isotope to study the oxidation behavior of lawrencium. Using solvent extraction techniques and working very rapidly, they extracted lawrencium ions from a buffered aqueous solution into an organic solvent -- completing each extraction in about 30 s.

Properties

Lawrencium behaves differently from dipositive nobelium and more like the tripositive elements earlier in the actinide series.

Sources: CRC Handbook of Chemistry and Physics and the American Chemical Society.

Last Updated: 12/19/97 *CST Information Services Team*

Periodic Table

Elements with their Symbol and Atomic Number

in alphabetical order

Symbol	Element	Atomic Number
Ac	Actinium	89
Al	Aluminum	13
Am	Americium	95
Sb	Antimony	51
Ar	Argon	18
As	Arsenic	33
At	Astatine	85
Ba	Barium	56
Bk	Berkelium	97
Be	Beryllium	4
Bi	Bismuth	83
B	Boron	5
Br	Bromine	35
Cd	Cadmium	48
Ca	Calcium	20
Cf	Californium	98
C	Carbon	6
Ce	Cerium	58
Cs	Cesium	55
Cl	Chlorine	17
Cr	Chromium	24
Co	Cobalt	27
Cu	Copper	29
Cm	Curium	96
Dy	Dysprosium	66
Es	Einsteinium	99
Er	Erbium	68
Eu	Europium	63

Symbol	Element	Atomic Number
Md	Mendelevium	101
Hg	Mercury	80
Mo	Molybdenum	42
Ns	Neilsborium	107
Nd	Neodymium	60
Ne	Neon	10
Np	Neptunium	93
Ni	Nickel	28
Nb	Niobium	41
N	Nitrogen	7
No	Nobelium	102
Os	Osmian	76
O	Oxygen	8
Pd	Palladium	46
P	Phosphorus	15
Pt	Platinum	78
Pu	Plutonium	94
Po	Polonium	84
K	Potassium	19
Pr	Praseodymium	59
Pm	Promethium	61
Pa	Protactinium	91
Ra	Radium	88
Rn	Radon	86
Re	Rhenium	75
Rh	Rhodium	45
Rb	Rubidium	37
Ru	Ruthenium	44

Fm	Fermium	100
F	Flourine	9
Fr	Francium	87
Gd	Gadolinium	64
Ga	Gallium	31
Ge	Germanium	32
Au	Gold	79
Hf	Hafnium	72
Ha	Hahnium	105
Hs	Hassium	108
Hi	Helium	2
Ho	Holmium	67
H	Hydrogen	1
In	Indium	49
I	Iodine	53
Ir	Iridium	77
Fe	Iron	26
Kr	Krypton	36
La	Lanthanum	57
Lr	Lawrencium	103
Pb	Lead	82
Li	Lithium	3
Lu	Lutetium	71
Mg	Magnesium	12
Mn	Manganese	25
Mt	Meitnerium	109

Rf	Rutherfordium	104
Sm	Samarium	62
Sc	Scandium	21
Sg	Seaborgium	106
Se	Selenium	34
Si	Silicon	14
Ag	Silver	47
Na	Sodium	11
Sr	Strontium	38
S	Sulfur	16
Ta	Tantalum	73
Tc	Technetium	43
Te	Tellurium	52
Tb	Terbium	65
Tl	Thalium	81
Th	Thorium	90
Tm	Thulium	69
Sn	Tin	50
Ti	Titanium	22
W	Tungsten	74
U	Uranium	92
V	Vanadium	23
Xi	Xenon	54
Yb	Ytterbium	70
Yb	Yttrium	39
Zn	Zinc	30
Zr	Zirconium	40

Los Alamos National Laboratory

Periodic Table of the Elements

Comments - Questions - Suggestions

Your Email Address:

Your Comments,
Questions or Suggestions :

Tell us | Clear

Back to the Elements

LOS ALAMOS NATIONAL LABORATORY
Operated by the University of California for the US Department of Energy

LANL | cstis@lanl.gov
For conditions of use, see Disclaimer

Periodic Table

Back

What is the Periodic Table of The Elements?

"It is a huge, efficient resource!"

The periodic table is the most important chemistry reference there is. It arranges all the known elements in an informative array. Elements are arranged left to right and top to bottom in order of increasing atomic number.. This order generally coincides with increasing atomic mass

The different rows of elements are called periods. The period number of an element signifies the highest energy level an electron in that element occupies (in the unexcited state). The number of elements in a period increases as one traverses down the periodic table because as the energy level of the atom increases, the number of energy sub-levels per energy level increases.

Using the data in the table scientists, students, and others that are familiar with the periodic table can extract information concerning individual elements. For instance, a scientist can use carbon's atomic mass mass to determine how many carbon atoms there are in a 1 kilogram block of carbon.

People also gain information from the periodic table by looking at how it is put together. By examining an element's position on the periodic table, one can infer the electron configuration. Elements that lie in the same column on the periodic table (called a "group") have identical valance electron configurations and consequently behave in a similar fashion chemically. For instance, all the group 18 elements are inert gases. The periodic table contains an enormous amount of important information. People familiar with how the table is put together can quickly determine a significant amount of information about an element, even if they have never heard of it.

LOS ALAMOS NATIONAL LABORATORY
Operated by the University of California for the US Department of Energy
LANL External View | www@lanl.gov | Help | Copyright © UC 1999 | Disclaimer

Periodic Table

Back

How to use the Periodic Table

When you open any file of an element in the periodic table, you will find a small table with some basic information about that element. Here's how you use that table:

5	Atomic Number
B	Atomic Symbol
10.81	Atomic Mass

- **Atomic Number**
  The number of protons in an atom defines what element it is. For example carbon atoms have six protons, hydrogen atoms have one, and oxygen atoms have eight. The number of protons in an atom is referred to as the atomic number of that element. The number of protons in an atom also determines the chemical behavior of the element.
- **Atomic Symbol:**
  The atomic symbol is one or two letters chosen to represent an element ("H" for "hydrogen," etc.). These symbols are used internationally. Typically, a symbol is the truncated name of the element or the truncated Latin name of the element. Click here for a list of the elements and their symbols.
- **Atomic Mass:**
  The atomic mass is the average mass of an element in atomic mass units ("amu"). Though individual atoms always have an integer number of amus, the atomic mass on the periodic table is stated as a decimal number because it is an average of the various isotopes of an element. Isotopes can have a weight either more or less than the average. The average number of neutrons for an element can be found by subtracting the number of protons (atomic number) from the atomic mass.
- **Electron Configuration:**
  The electron configuration is the orbital description of the locations of the electrons in an unexcited atom. Using principles of physics, chemists can predict how atoms will react based upon the electron configuration. They can predict properties such as stability, boiling point, and conductivity. Typically, only the outermost electron shells matter in chemistry, so we truncate the inner electron shell notation by replacing the long-hand orbital description with the symbol for a noble gas in brackets. This method of notation vastly simplifies the description for large molecules.
  Example: The electron configuration for Be is $1s^2 2s^2$*, but we write* $[He]2s^2$ *where [He] is equivalent to all the electron orbitals in the helium atom. The Letters, s, p, d, and f designate the shape of the orbitals and the superscript gives the number of electrons in that orbital.*

Further explanation of the atom may be found Here.

LOS ALAMOS NATIONAL LABORATORY
Operated by the University of California for the US Department of Energy
LANL External View | www@lanl.gov | Help | Copyright © UC 1999 | Disclaimer

Periodic Table

Back

Chemistry in a Nutshell

- **The Atom**
  All macroscopic matter is made out of many tiny particles called atoms. The study of how these atoms interact is called *Chemistry*.
- **Subatomic Particles**
  The three particles that make up atoms are protons, neutrons, and electrons. Protons and neutrons are heavier than electrons and reside in the "nucleus," which is the center of the atom. Protons have a positive electrical charge, and neutrons have no electrical charge. Electrons are extremely lightweight and are negatively charged. They exist in a cloud that surrounds the atom. The electron cloud has a radius 10,000 times greater than the nucleus.
- **The Nucleus**
  The nucleus of an atom is made up of protons and neutrons in a cluster. Virtually all the mass of the atom resides in the nucleus. The nucleus is held together by the tight pull of what is known to chemists and physicists as the "strong force." This force between the protons and neutrons overcomes the repulsive electrical force that would, according to the rules of electricity, push the protons apart otherwise.
- **Electrons**
  The electron is the lightweight particle that "orbits" outside of the atomic nucleus. Chemical bonding is essentially the interaction of electrons from one atom with the electrons of another atom. The magnitude of the charge on an electron is equal to the charge on a proton. Electrons surround the atom in pathways called orbitals. The inner orbitals surrounding the atom are spherical but the outer orbitals are much more complicated.
- **Chemical Bonding**
  Chemically bonding occurs when two particles can exchange or combine their outer electrons in such a way that is *energetically favorable*. An energetically favorable state can be seen as analogous to the way a dropped rock has a natural tendency to fall to the floor. When two atoms are close to each other and their electrons are of the correct type, it is more energetically favorable for them to come together and share electrons (become "bonded") than it is for them to exist as individual, separate atoms. When the bond occurs, the atoms become a compound. Like the rock falling to the floor, they "fall" together naturally.

LOS ALAMOS NATIONAL LABORATORY

Operated by the University of California for the US Department of Energy
LANL External View | www@lanl.gov | Help | Copyright © UC 1999 | Disclaimer

Periodic Table

Back

Mendeleev noted patterns in the combining ratios of elements

- The Elements Lithium (Li), Sodium (Na), and Potassium (K) all formed oxides in the ratio of two atoms per oxygen atom: R_2O
- The Elements Beryllium (Be), Magnesium (Mg), and Calcium (Ca) all formed oxides in the ratio of one atom per oxygen atom: RO
- Boron (B) and Aluminum (Al) formed R_2O_3
- Carbon (C) and Silicon (Si) formed RO_2

Recognizing the patterns of combining ratios or "valency", Mendeleev created a table organized by placing elements with similar combining ratios in the same group. He arranged the elements within a group in order of their atomic mass.

TABELLE II

REIHEN	GRUPPE I. — R^2O	GRUPPE II. — RO	GRUPPE III. — R^2O^3	GRUPPE IV. RH^4 RO^2	GRUPPE V. RH^3 R^2O^5	GRUPPE VI. RH^2 RO^3	GRUPPE VII. RH R^2O^7	GRUPPE VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199.
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	— — — —

Figure 2.5 Dmitri Mendeleev's 1872 periodic table. The spaces marked with blank lines represent elements that Mendeleev deduced existed but were unknown at the time, so he left places for them in the table. The symbols at the top of the columns (e.g., R^2O and RH^4) are molecular formulas written in the style of the 19th century.

In 1869, the Russian chemist Mendeleev noted that the repeating patterns of behavior could be arranged in a sequence of elements giving rise to the "Periodic Table" of the elements.

Special thanks to Dr. Paul Karol's "Intro to Modern Chemistry" for providing much of the information on this page.

LANL External View | www@lanl.gov | Help | Copyright © UC 1999 | Disclaimer

About this Resource

Originally this resource, the Periodic Table, was created by Robert Husted at Los Alamos National Laboratory during his time as a Graduate Research Assistant.

The Periodic Table that you are currently viewing was inherited by the Chemistry Division from the Computer Division who provided the laboratory some of the internets first web sites..

This page is being offered as a public service to all those wanting to become more educated in the field of Chemistry. This is an imperfect resource with ambiguities and errors. If you notice something incorrect about our web page we would be grateful for any pointers. We welcome any suggestions or comment at cstsupport@lanl.gov.

Good Luck, and Good Chemistry.

Chemistry Division

Computer Information Services Team